EXPERIMENTAL SCIENCE

FREWIN

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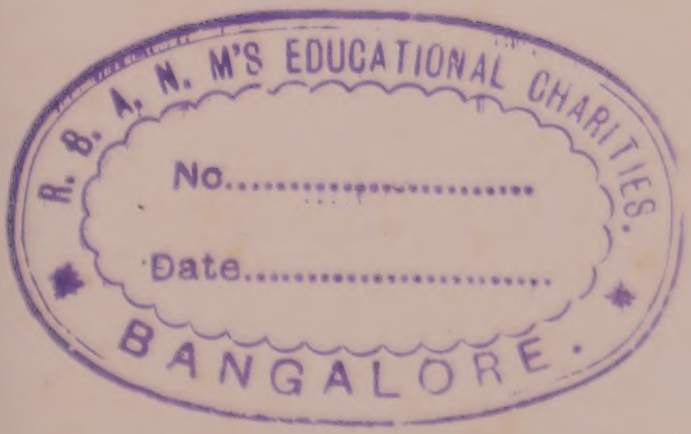
A New EXPERIMENTAL SCIENCE

BY

J. G. FREWIN

PART II



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'In Nature's infinite book of secrecy,
a little can I read.'

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PREFACE

DURING recent years the content of the schemes of work in school science has undergone considerable modification, especially in the earlier portions of the course. Experience has shown that when much time has been devoted to measurement, however desirable practice in such exercises may be, the younger pupils lose interest and the laboratory soon ceases to be an attraction. Boys and girls are naturally curious and delighted to find out about things around them. Efforts are being made in schools to develop along this natural line. The numerical conundrums that inevitably followed experiments, as if science had been brought into schools to provide practical material for Arithmetic, have had their day. 'Science is measurement', but measurement is of little value to the pupils until they have made acquaintance with facts and principles through experiment, and see why measurements are required.

The writer is much indebted to the science staffs of many schools for their readiness to criticize and suggest, and the scheme on which this book is based embodies much that has been gathered from these schools; also to the General Electric Company for the Electric Lifting Magnet.

The order of the main divisions of the book is unusual. The object aimed at, before all else, has been to interest the pupils in the subject, and the method adopted has been to avoid monotony by changing the type of work each Term.

In conclusion, the scheme is not an untried one. It represents what may be termed the standard course for a large number of schools, and is put into book form in response to oft-repeated request.

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I

ELEMENTARY CHEMISTRY OF WATER

WATER

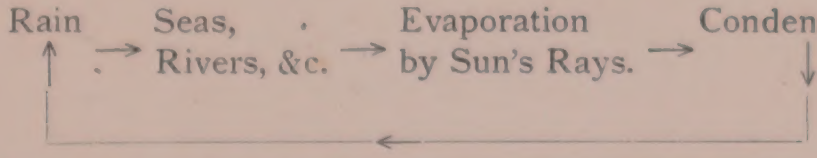
WATER is very abundantly distributed. A glance at the map of the world will show that three-fourths of the earth's surface is covered with water.

Water is as necessary for life as air. Men have voluntarily undertaken long fasts, living without food for a month, or even longer, but they were compelled to drink water. Our bodies are about 70 per cent. water, i.e. of the 140 lb. that a pupil in the senior school weighs, 98 lb. are water. The water content of plants and animals varies considerably.

These figures, which are quite general, are interesting:

Plants	. . .	60–90 per cent.	water.
Fish	. . .	80 ,,	,,
Animals	. . .	70 ,,	,,

We have already referred to the Carbon cycle (Bk. I. 70). Think of the water cycle.



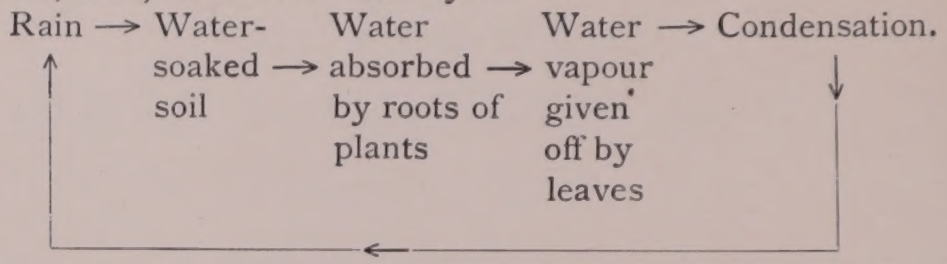
Plants taken from the soil, or flowers cut from the plant, soon wither.

Experiment 1. Place a bunch of leaves under a bell-jar and note the beads of moisture, i.e. water, that collect on the inside surface of the jar. Account for them.

Experiment 2. Take a bunch of weeds and place them in a gas jar containing water so that the roots are submerged. Add two or three drops of oil to the water to form a surface film and

so prevent evaporation. Mark the level of the water with gummed paper. Leave for about a week. Note the change in water-level.

Here, then, is another water cycle.



As in the case of **Air**, we shall make our study of water under the two heads—Chemistry and Physics.

We have learned that Chemistry is concerned with the composition of substances. The chemist experiments in an endeavour to break up a substance until he is confronted with materials that will not give way. These, as we have seen, are called elements. Elements are the simplest substances. From mercury oxide we obtained Mercury and Oxygen. Mercury and Oxygen are elements. Is water an element? If not, we should be able to analyse water into its elements and synthesize it from its elements.

Experiment 3. Make a little magnesium oxide by burning a piece of Magnesium ribbon in the air. Note the colour of the oxide. Shake it up with water in a test-tube and test the solution with litmus paper.

Magnesium + Oxygen → Magnesium oxide.
(metal) (gas) (ash)

Magnesium oxide: A greyish-white solid.
Sparingly soluble in water.
Solution in water is alkaline.

Magnesium oxide is quite unlike Magnesium or Oxygen.

Recall:

Blue copper sulphate crystals containing Copper, but quite unlike Copper.

Red oxide of Mercury containing Mercury and Oxygen, but quite unlike Mercury or Oxygen.

Iron oxide (red rust) containing Iron and Oxygen.

Carbon dioxide (colourless gas) containing Carbon and Oxygen.

When two or more elements are combined together to form an entirely new substance quite unlike the constituents, the

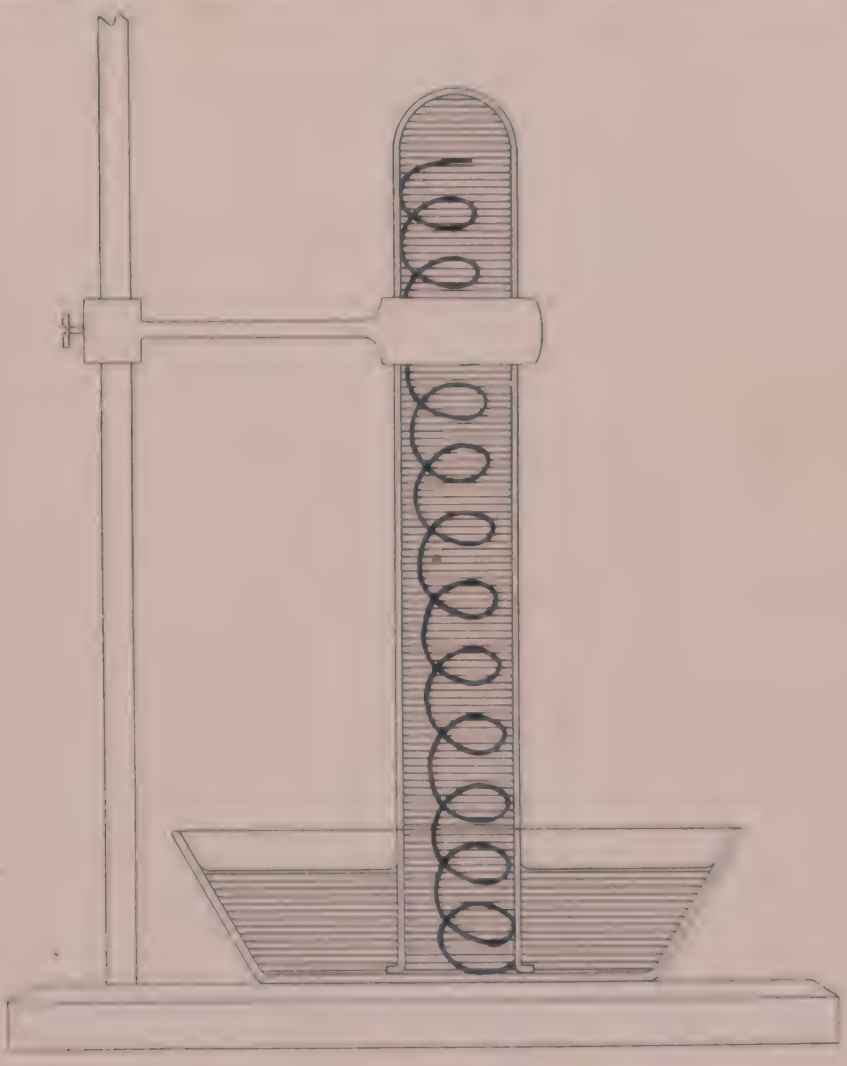


Fig. 1.

new substance is said to be a **Chemical Compound**. Magnesium oxide is a chemical compound of Magnesium and Oxygen. Copper sulphate is a chemical compound of Copper, Sulphur, and Oxygen, and so on.

Experiment 4. Clean the surface of about two feet of Magnesium ribbon with emery paper. What do you remove? Coil

it tightly round your pencil and push the coil formed into a test-tube. Fill the test-tube with water and invert it over a dish of water. See that there is no air-bubble at the top of the test-tube. Leave the apparatus set up until next lesson. (Fig. 1.)

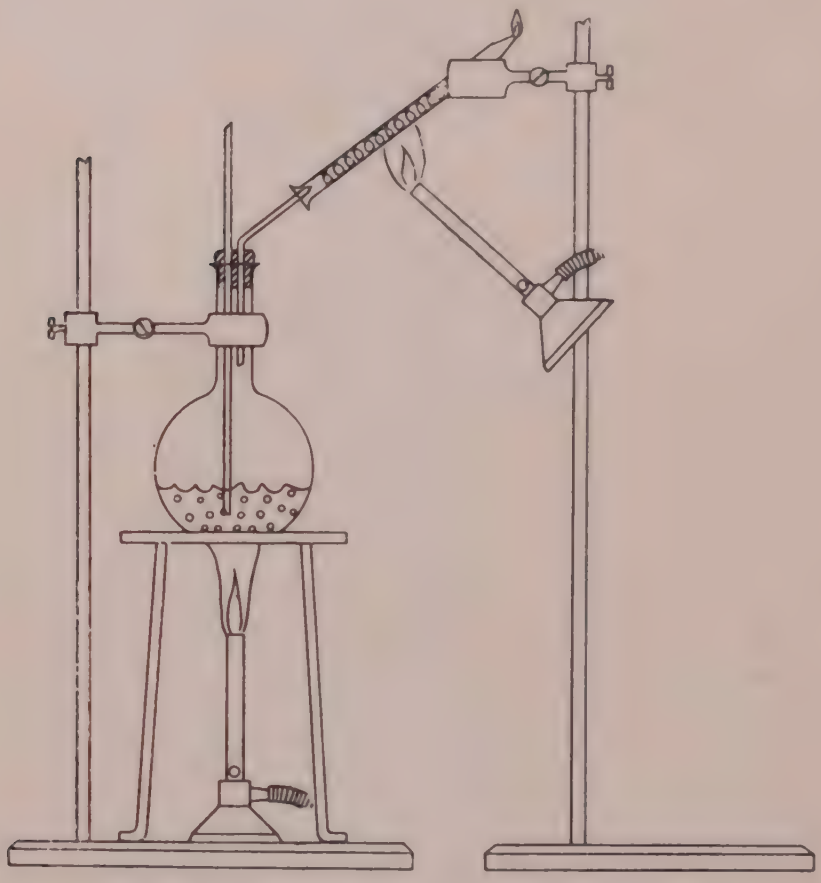


Fig. 2.

Experiment 5 (Demonstration). Take a piece of hard glass tubing ($\frac{3}{8}$ in. diameter) six inches long and draw out a jet at one end.

Wind a coil of Magnesium ribbon round the wire of a deflagrating spoon and push the coil into the hard glass tube.

Fit up a flask with safety tube and exit tube as shown in the figure. Half fill with water and boil. When steam is issuing freely, heat the hard glass tube containing Magnesium, and then direct the current of steam into the heated Magnesium coil. (Fig. 2.)

As soon as the action starts, hold a lighted match to the jet.

Note: (I) The Magnesium glows brightly in the steam, as it did in the air that was allowed to enter the crucible in a former experiment. (Bk. I, p. 55.)

(II) Some of the Magnesium is converted into magnesium oxide. Compare its properties with those of the magnesium oxide obtained by burning Magnesium in the air.

(III) **An inflammable gas burns at the jet.**

This inflammable gas was named **Hydrogen** by Lavoisier in 1783. (Gk. *hudor*, water, as in hydrant—a water-pipe; hydraulic, operated by water.)

Oxygen→The acid producer (Bk. I. 65).

Nitrogen→The nitre producer.

Hydrogen→The water producer.

The names Nitrogen and Hydrogen are sometimes confused. *Remember* Hydrogen is Highly Inflammable.

Nitrogen is Not Inflammable.

Look at these two statements:

Magnesium + Oxygen→Magnesium oxide.

Magnesium + Steam→Magnesium oxide + Hydrogen.

It is not difficult to see that the Oxygen and Hydrogen of the right-hand side must have been supplied by the steam of the left-hand side, for Magnesium is an element.

Experiment 6. Examine the apparatus set aside (Exp. 4). Note the gas collected in the test-tube and the change in the appearance of the Magnesium. Place the thumb over the mouth of the test-tube, before lifting it out of the water, to prevent the escape of the gas. Invert the test-tube. Ask your neighbour to be ready with a lighted match, and let him hold the flame at the mouth of the test-tube the moment you remove your thumb. If you do this carefully you should see the gas take fire.

Put pieces of red and blue litmus in the water. Is the solution acid or alkaline?

Magnesium +	Water →	Magnesium oxide +	Hydrogen.
(bright metallic element)	(neutral liquid)	(Grey powder, solution in water alkaline)	(inflammable gas)

Is water an element?

Water is composed of Hydrogen and Oxygen and is therefore **hydrogen oxide.**

Using chemical names only, we write

Magnesium + Hydrogen oxide → Magnesium oxide + Hydrogen.

Water is quite unlike the two elements Oxygen and Hydrogen. A fire can be extinguished by pouring water on it. What would happen if the hose delivered Hydrogen, or Oxygen, or both, instead of water?

Water is, therefore, a Chemical Compound. It is no more like Oxygen and Hydrogen than the white ash magnesium oxide is like Oxygen and Magnesium.

When the substances left at the end of an experiment are quite unlike the substances used at the beginning, as in this and many other cases you have met, they are said to be the result of **Chemical Action.** Chemical action produces **Permanent Change.**

Notice that as soon as we cause permanent changes in the material with which we are working we study it from the chemist's standpoint. In many experiments, e.g. freezing and boiling of water, compression of air, expansion of Mercury on heating, there was no permanent change in the material used. These experiments came under the heading of Physics. To decide whether any action is a Chemical or Physical Action ask the question: 'Has the material employed changed, really and permanently?'

Several other metals will combine with the Oxygen in water and liberate the Hydrogen.

Experiment 7 (Demonstration). Fit up the apparatus as shown in figure (Fig. 3), and pass steam over red-hot iron nails, contained in a hard glass tube (about 1 in. diam.).

Collect the gas in a test-tube over water and ignite it.

Iron + Water → Iron oxide + Hydrogen.

Experiment 8. Invert a test-tube of water over a small dish of water. Place a few shavings of Calcium under the test-tube and test the gas evolved. (Calcium is an element.)

Calcium + Water → Calcium oxide + Hydrogen.

Filter the water used in this experiment and blow through a glass tube into the clear liquid. What happens? (Bk. I, p. 67.)

We have already learned the test for Carbonic Acid Gas. When Carbonic Acid Gas is bubbled through lime water, the solution becomes chalky or milky; so you will not be surprised to hear that lime and this calcium oxide are one and the same thing. Which of these names is the chemical name? Is calcium oxide a Chemical Compound? Why?

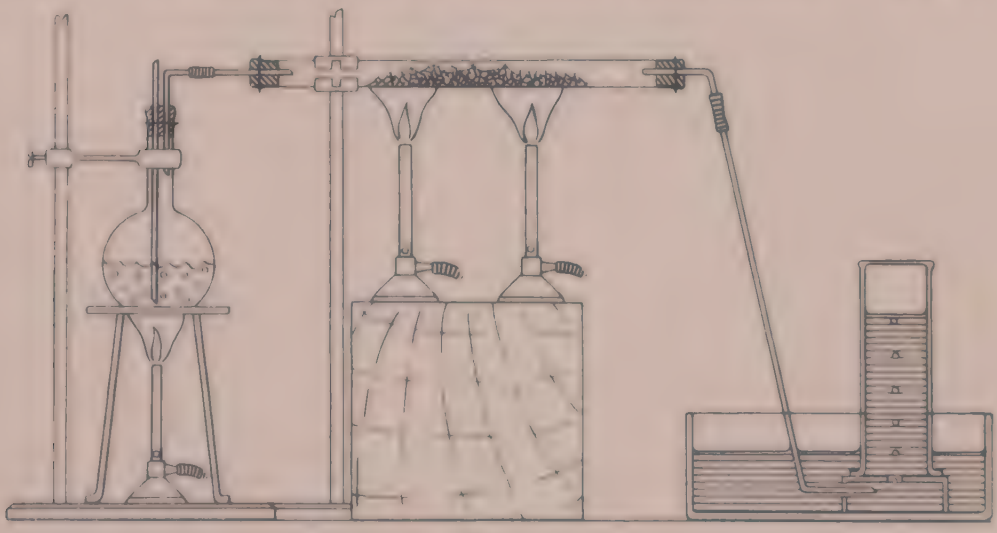


Fig. 3.

Calcium	+	Hydrogen oxide	⟶	Calcium oxide	+	Hydrogen.
(metal, element)		(liquid compound)		(solid compound)		(gas, element)

Is this a Chemical Action?

Hydrogen is a constituent of many substances, e.g. Sulphuric Acid.

Blue copper sulphate crystals, as we have seen, contain the element Copper. The Copper can be displaced by Iron or Zinc. (Bk. I, p. 47.)

Experiment 9. Dissolve a few crystals of copper sulphate in warm water in an evaporating basin. You will find it takes less time if you first grind the crystals to powder in a mortar. When eating a hard sweet we endeavour to crush it to powder with the teeth. Think out an explanation.

Add small pieces of granulated Zinc and note the change in the

colour of the solution. Continue until the blue colour has disappeared and a little Zinc is left undissolved.

Fit up a filter funnel with filter paper and pour the contents of the basin into it. Catch the clear filtrate, i.e. the liquid that passes through the filter paper.

What is on the filter paper? Account for it.

Evaporate the filtrate until the crystals form on a glass rod, dipped into it and allowed to cool. A sand bath should be used. (Bk. I. 25, 26.)

Set aside to crystallize.

The white crystals are zinc sulphate.

Zinc + Copper sulphate → Zinc sulphate + Copper.

We shall now perform a similar experiment with Sulphuric Acid.

Sulphuric Acid is a liquid that must be handled with caution. It destroys cloth, so be careful not to spill drops on your clothes or on the table. Should any be spilt, wipe it up with a damp cloth immediately.

Experiment 10. Pour a little dilute Sulphuric Acid into an evaporating basin. Add small pieces of granulated Zinc until effervescence ceases, leaving some pieces of Zinc undissolved. What does effervescence indicate? Fit up a filter funnel with filter paper and pour the contents of the basin into it. Catch the clear filtrate. Evaporate it until crystals form on a piece of glass rod dipped into it and allowed to cool. A sand bath or water bath should be used. Set it aside to crystallize. The crystals are zinc sulphate.

Zinc + Sulphuric Acid → Zinc sulphate + . . . a gas.

We have not studied this action completely. The effervescence showed that a gas was being evolved. What is this gas?

Experiment 11. Take a test-tube and about one-third fill it with dilute Sulphuric Acid. Add granulated Zinc and test the gas by holding a lighted match to the mouth of the test-tube.

The gas will ignite, often explosively, and burn with an almost invisible blue flame. This is the test for Hydrogen.

Zinc + Sulphuric Acid → Zinc sulphate + Hydrogen.

Evidently, here, Hydrogen takes the place of the Copper in the experiment with Zinc and Copper sulphate. So we may write

Zinc + Hydrogen sulphate → Zinc sulphate + Hydrogen.

By means of these experiments we have been able to determine the chemical name for Sulphuric Acid, viz. Hydrogen sulphate. What are the three elements of which Sulphuric Acid is composed?

Zinc is not the only metal that will displace the Hydrogen from Sulphuric Acid.

Experiment 12. Use a test-tube as in the last experiment and try the action of Iron, Copper, Magnesium, Lead, on dilute Sulphuric Acid. If there is effervescence, test for Hydrogen.

Finish these statements:

Zinc + Hydrogen sulphate → Zinc sulphate + Hydrogen.
Iron + Hydrogen sulphate →
Copper + Hydrogen sulphate →
Magnesium + Hydrogen sulphate →
Lead + Hydrogen sulphate →

Where there is no effervescence, write 'No action'.

We shall now examine the properties of this gas. The most convenient way of preparing it in the laboratory is in the apparatus shown in Fig. 4.

The double-necked bottle is known as a Woulfe's bottle. It was first described by Woulfe in 1784. Note how often

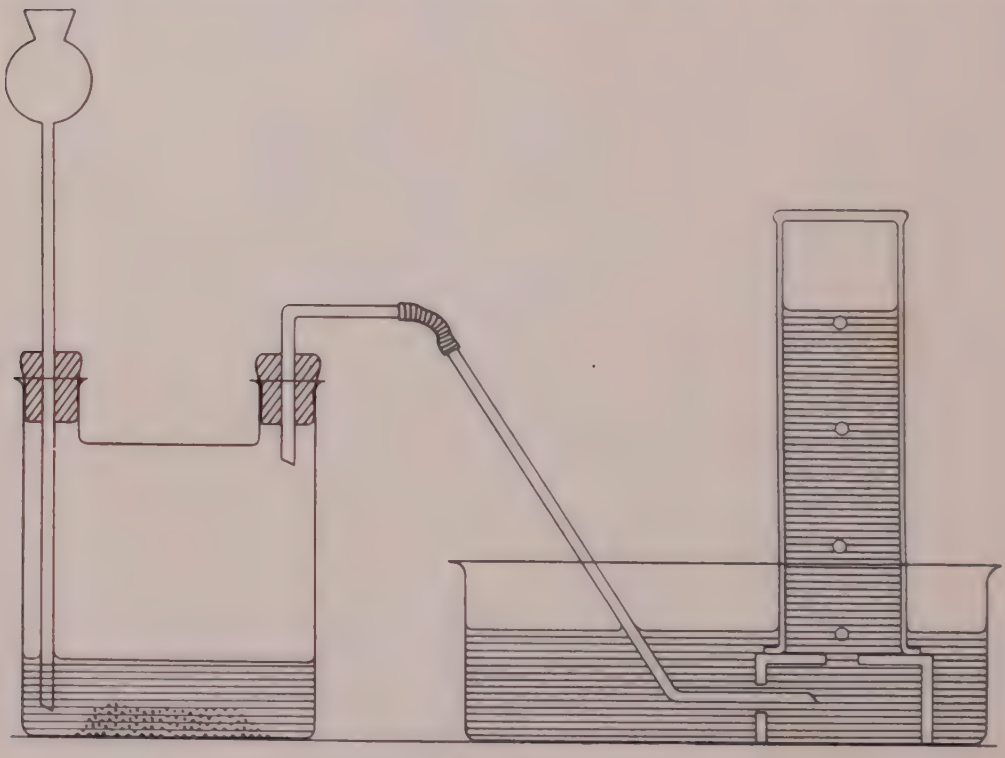


Fig. 4.

the latter half of the eighteenth century is mentioned. (A flask fitted with a two-hole stopper will do, but is very liable to be broken when the Zinc is put in, unless great care is taken. Hold the flask almost horizontally.)

Experiment 13. Put some granulated Zinc into the Woulfe's bottle. Fit up the apparatus (Fig. 4). Be sure that the end of the thistle funnel is close to the bottom of the bottle or the gas will escape up it, into the air, instead of passing through the delivery tube to the gas jar over the pneumatic trough.

Pour dilute Sulphuric Acid down the thistle funnel. Collect

six jars of the gas, keeping the first one apart from the others. There must not be any naked lights near the apparatus. Look back at Experiment 11.

Cover each jar, when it is filled with Hydrogen, with a glass plate and stand the jars in a row on the table. Dismantle the apparatus before experimenting with the gas.

Hydrogen is a colourless, odourless gas.

Jar 1. Light a match or taper, remove the glass plate from the *first* jar that was collected and immediately hold the taper to the mouth of the jar. Point the mouth of the jar *away* from you.

Jar 2. Treat this jar as you treated the first and compare the results.

Which jar contains the Hydrogen? What gases are mixed with the Hydrogen in the first jar?

Hydrogen burns with a blue flame. When mixed with air it ignites with a sharp explosion.

Which of the gases in the air, Nitrogen or Oxygen, causes the Hydrogen to burn explosively? Give reasons for your answer.

Jars 3 and 4. Hold jar 3 mouth upwards and jar 4 mouth downwards. Remove the glass plates for about ten seconds. Replace. Test the gas in each jar with a lighted taper. From which jar has the gas escaped? Can you suggest the cause of this?

Jar 5. Pour the Hydrogen from this jar into another jar containing air and prove (1) that the Hydrogen has left jar 5, and (2) that the Hydrogen has been transferred to the other jar.

Think out how you will do this and ask whether your method is correct before trying it.

A balloon or soap-bubbles filled with Hydrogen will ascend to the ceiling. This should be shown as a demonstration experiment. Remove the thistle funnel and close the neck with an ordinary cork. Fasten a rubber balloon to the delivery tube. When the balloon is filled, tie the mouth and set it free. Why was the thistle funnel removed?

Jar 6. Hold the jar of Hydrogen mouth downwards (why?). Take a long taper, light it; remove the plate from the jar and push the lighted taper right up into the jar.

Does Hydrogen support Combustion? Withdraw the taper slowly. What happens? Why? Is Hydrogen soluble in water? How does the method used in collecting the gas suggest the answer to this question?

Write out the properties of **Hydrogen.**

Experiment 14 (Demonstration). Fit a jet instead of the delivery tube to the Woulfe's bottle. Prepare Hydrogen as in the previous experiment. Make sure that all fittings are tight. When the effervescence is vigorous, cover the Woulfe's bottle with a duster to lessen the risk of scattered glass should an explosion occur. Hold an inverted test-tube over the jet for a few seconds, remove to a distance and apply a light. If the gas ignites *quietly*, light the Hydrogen issuing from the jet. Use the *quietly* burning Hydrogen to light it. Bring a flask filled with cold water, but dried and polished on the outside, over the flame. (Fig. 5.) Watch the surface of the flask.

Priestley was the first to notice that when Hydrogen burns moisture is produced.

What experiments could you perform to show that the substance is Water? (Bk. I, pp. 39, 43.)

What is formed when Magnesium burns in air?

What is formed when Carbon burns in air?

What is formed when Hydrogen burns in air?

This experiment has shown, by synthesis, that water is hydrogen oxide.

In 1800 two scientists, who were experimenting with an electric battery, put a drop of water in contact with the two wires and noticed that bubbles of gas were formed round them. This led them to put the two wires into a beaker of water and collect the bubbles of gas from each. They found that the gas evolved at one wire was Oxygen and the gas evolved at the other wire was Hydrogen. How would you test for these gases?

This experiment is a very interesting one, and is known as the Electrolysis of Water. What does Electrolysis mean? Compare Analysis, Synthesis.

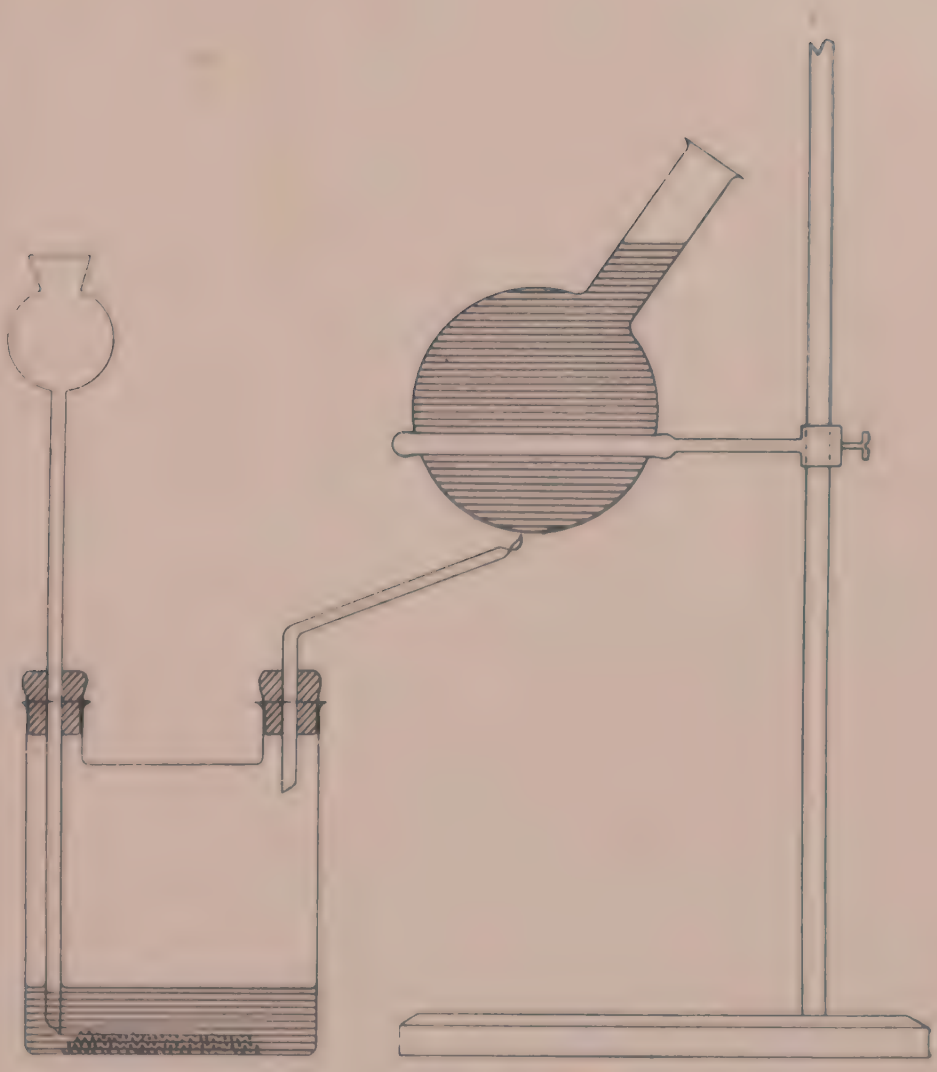


FIG. 5.

Experiment 15 (Demonstration). Half fill a large evaporating basin with water—slightly acidulated with sulphuric acid—and stand two inverted test-tubes of acidulated water in the basin. Two copper wires, rubber insulated or enamelled, to each of which a small piece of platinum foil is fixed, are attached to a battery (four bichromate cells). The wires are bent so as to bring the platinum within the test-tubes. See Fig. 6.

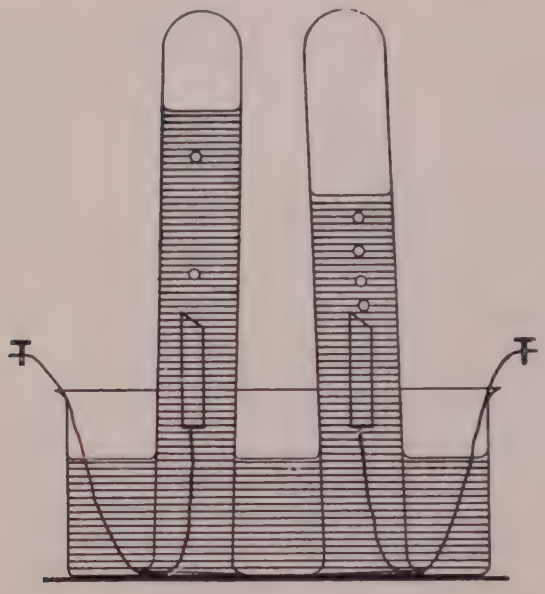


Fig. 6.

Note the amount of gas in each test-tube and show by the tests that the larger quantity is Hydrogen, and the smaller quantity is Oxygen. When water is decomposed by electricity, the Hydrogen liberated is twice the volume of Oxygen liberated.

Chemists write the symbol H_2O for water. You will notice that the chemist is not quite satisfied when he has found out the elements of which a substance is composed; his next inquiry is directed to finding information about the quantities of the elements discovered. H_2O carries us a step further than 'hydrogen oxide'.

Coal-gas. Purified coal-gas may contain as much as 50 per cent. of Hydrogen. It will therefore behave much like Hydrogen, and answers quite well for some experiments.

Experiment 16 (Demonstration). Fit a glass tube to the Bunsen gas-tubing and blow soap-bubbles with coal-gas. Notice what happens when the bubbles are freed from the soap solution. Put a lighted match to one.

Experiment 17 (Demonstration). Take a cocoa or syrup tin and pierce one hole in the centre of the lid and one in the bottom. The holes should be about $\frac{3}{16}$ in. in diameter. Place the tin on a tripod stand and pass a stream of coal-gas through the lower

hole into the tin. Air will escape from the upper hole. In about thirty seconds remove the tube and at once apply a light to the hole in the lid. Coal-gas, being less dense than air, issues from this hole and air comes in from below to take its place. In a short time the gas remaining in the tin will be thoroughly mixed with air, containing Oxygen; the tiny flame will pass through the hole and the explosion will probably blow the lid of the tin to the ceiling.

This experiment should warn us not to look for escaping gas with a lighted match, and never to turn on the gas in a gas-oven unless the oven door is wide open.

When coal-gas burns heat is produced. In this experiment the gas and air are thoroughly mixed, so that the burning of the gas in the tin is almost instantaneous. The water and carbon oxide produced are therefore very hot, and, owing to expansion, occupy a large volume.

Coal-gas (Carbon combined with Hydrogen) + Oxygen (in the Air) ⟶ Carbon oxide + Hydrogen oxide.

Petrol is also composed of Carbon and Hydrogen. In the cylinders of a motor, petrol vapour is mixed with air and ignited by an electric spark.

Experiment 18. Open the air holes of the Bunsen burner, turn the gas half on, and attempt to light the gas. This often results in 'striking back'. You should be able to explain this. It frequently happens where incandescent lighting, gas-fires, gas-rings, gas-ovens, are used. In all these we are employing Bunsen burners.

When you light any form of Bunsen burner wait two or three seconds for the gas to sweep out the air in the pipe.

REDUCTION

Experiment 19. Take a piece of combustion tubing (diameter $\frac{3}{16}$ in. or $\frac{1}{2}$ in.) about 8 in. long and fit it up as in the sketch.

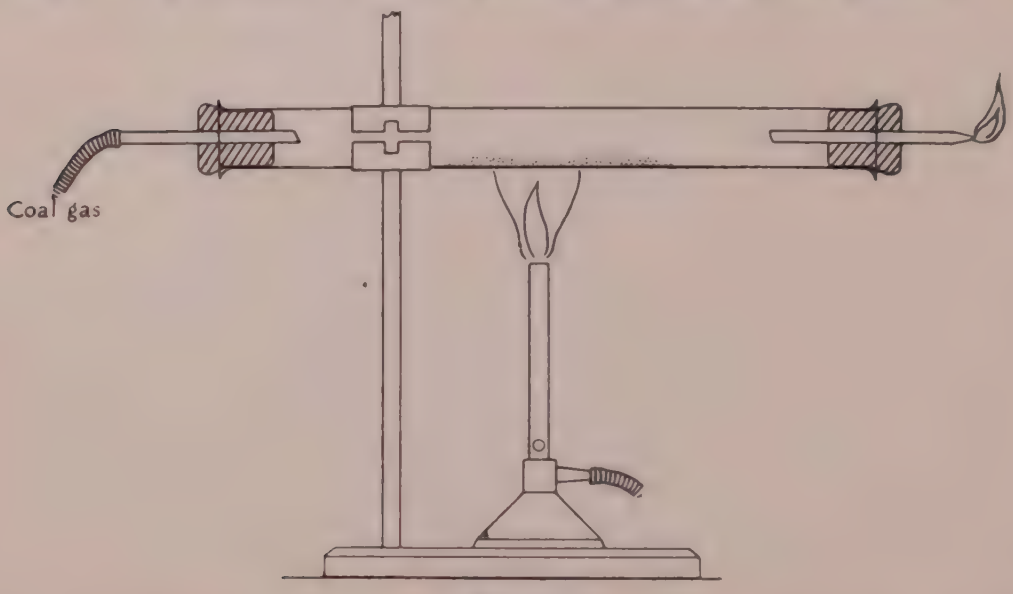


FIG. 7.

(*a*) Spread a little copper oxide (wire) in the middle of the tube. Heat the copper oxide and allow coal-gas to pass *slowly* over it. Light the issuing gas. The flame should be quite small. The Hydrogen in the coal-gas acts on the copper oxide, taking away the Oxygen to form water and leaving copper behind. Look for the drops of water formed. Is this a Chemical Action?

Hydrogen + Copper oxide ⟶ Hydrogen oxide + Copper.
(of coal-gas) (or water)

(*b*) Take out the Copper and put a little red lead, i. e. lead oxide, in its place. Pass coal-gas over the heated lead oxide, and write down the results as in (*a*).

Two tests for lead: (1) lead beads are easily flattened.
(2) lead beads make a black mark on paper.

When elements combine with oxygen they are said to be **OXIDIZED.** Iron left out of doors becomes **oxidized.** When **Oxygen** is removed from an **Oxide** the latter is said to be **REDUCED.** Hydrogen is a very powerful reducing agent. In

the experiments just completed hydrogen reduced the copper oxide and lead oxide to Copper and Lead respectively. What happened to the Hydrogen? Having seen the experiment with coal-gas and copper oxide, you will be able to follow a very important experiment carried out by **Dumas** in 1842. (Fig. 8, page 23.)

(1) Woulfe's bottle containing Zinc and Sulphuric Acid to furnish Hydrogen.

(2) (3) (4) (5) (6) (7) U-tubes containing chemicals to absorb water and other impurities in the Hydrogen.

(6) was immersed in a freezing mixture.

(8) Bulb containing copper oxide.

(9) Bulb to catch condensed water.

(10) (11) (12) (13) U-tubes containing chemicals to absorb all traces of water.

He passed *pure dry* hydrogen over heated *pure dry* copper oxide. He took every known precaution to be sure that the substances were really pure and dry. (Why?) The water formed in the experiment was collected and weighed. The bulb containing the copper oxide was weighed, before and after the experiment, to find the weight of the Oxygen given up. He performed the experiment nineteen times, and found the average of his results to be as follows:

Wt. of water produced	49·76 gm.
Wt. of Oxygen lost by copper oxide	44·22 gm.
The difference gives the weight of Hydrogen that has combined with this Oxygen (49·76 − 44·22 gm.)	5·54 gm. Hydrogen.

Copper oxide + Hydrogen → Water + Copper.

218·3 gm. 5·54 gm. 49·76 gm. 174·1 gm.

5·54 gm. Hydrogen. 44·22 gm. Oxygen.

44·22 is almost exactly eight times 5·54

$5{\cdot}54 \times 8 = 44{\cdot}32$.

He concluded, therefore, that water is composed of 8 parts by weight of Oxygen combined with 1 part by weight of Hydrogen. So 9 lb. of water contain 8 lb. of Oxygen and 1 lb. of Hydrogen.

Now the volume of Oxygen, as we have seen by electrolysis, is half the volume of Hydrogen, and Dumas showed that it weighed eight times as much. If then we had the same volume of Oxygen it would weigh sixteen times as much.

From these two experiments we are able to conclude that Oxygen is sixteen times as dense as Hydrogen, and since Oxygen and the air are nearly of equal density it is quite easy to see why Hydrogen is used for filling balloons and airships.

1 c.cm. Hydrogen weighs 0·00009 gm.
1 c.cm. Oxygen weighs 0·00144 gm.
1 c.cm. Air weighs 0·00124 gm.

These numbers are too small to think about. It will help you to imagine a cubic yard, i. e. the volume of a cubical box, each face 1 sq. yd. In a corner of the room mark on the floor, with chalk, a sq. yd.; on each of the two walls of the corner also mark a sq. yd. The cubical box (1 cu. yd.) would fit into the marked space.

1 cu. yd. Hydrogen weighs about $2\frac{1}{2}$ oz.
1 cu. yd. Oxygen weighs about $2\frac{1}{2}$ lb.
1 cu. yd. Air weighs about 2 lb.

Chemists have experimented in several quite different ways to find the proportions, by *weight*, in which the Oxygen and Hydrogen combine to form water, and the results always agree closely with those found by Dumas, viz. 8 : 1. Very many different compounds have been tried and many different sources have been sought for one and the same compound, but the results are always the same. This agreement is expressed as a Law of Chemistry.

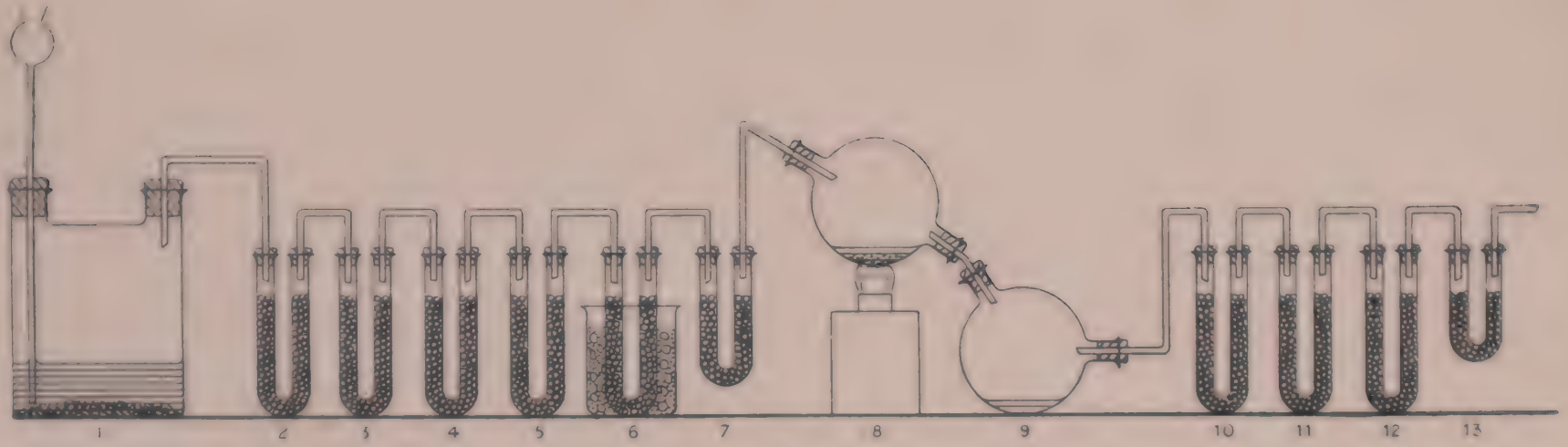


FIG. 8.

THE LAW OF CONSTANT COMPOSITION

A Chemical Compound always contains the elements of which it is composed united in the same proportions.

Jean Rey, you will remember, realized this in his experiment. (Bk. I, p. 56.)

Another chemist at the beginning of the nineteenth century said, 'Nature never produces a compound other than balance in hand'. He instanced common salt, which, no matter what part of the world is its source, has always the same composition.

Magnesium was converted to magnesium oxide in Experiment 39, Bk. I. The result of that experiment when carefully carried out is always the same. One gm. magnesium oxide is composed of 0·6 gm. Magnesium and 0·4 gm. Oxygen.

Magnesium	. .	60 per cent.
Oxygen .	. .	40 ,,
		100

Experiment 20. What is the proportion of Oxygen in potassium chlorate?

Take a hard test-tube and put in it a little manganese dioxide. Weigh the test-tube and manganese dioxide. Add about 0·5 gm. of potassium chlorate and weigh again. Shake the two substances together to mix them. Heat, slowly at first, and then strongly, for about five minutes. Allow to cool. Weigh and calculate the loss per cent.

Record:

Weight of test-tube + manganese dioxide	=	gm.
Weight of test-tube + manganese dioxide + potassium chlorate	=	gm.
Weight of potassium chlorate	=	gm.
Weight of test-tube and contents after Oxygen has been driven off	=	gm.
Weight of Oxygen driven off	=	gm

Why was the manganese dioxide put in? What is a Catalyst? (Bk. I, p. 61.) What weight of Oxygen would be given off on heating 100 gm. potassium chlorate? A carefully conducted experiment always gives 39·1 per cent. of Oxygen.

Experiment 21. Grind a very little copper sulphate to powder. Tip it into a hard test-tube and heat, at first cautiously, and then strongly, until it has quite lost its blue colour. What do you see on the side of the test-tube? What has been driven out of the copper sulphate crystals by heating them? Add two or three drops of water to some of the white powder and note the result. Reserve a small quantity.

Copper sulphate is really a white non-crystalline substance. When, however, it is dissolved in water it forms a blue solution from which it is quite easy to obtain large blue crystals. These crystals contain water, for (1) they weigh more than the white non-crystalline copper sulphate that was originally dissolved, and (2) water is given off on heating them.

Experiment 22. Take some small crystals of Washing Soda and heat them. Note that a liquid condenses on the side of the test-tube. Add a little of the white powder reserved from the last experiment to this liquid. What do you conclude from the result?

Crystalline substances that contain water as a necessary part of the crystals and which, on heating, give up the water and become non-crystalline are said to contain 'Water of Crystallization'. When this is driven off, the non-crystalline powder is termed *anhydrous* (Gk. *an*, without; *hudor*, water).

When we buy a pound of washing soda, the crystals contain ten ounces of water of crystallization.

Experiment 23. Verify the last statement. Use a crucible and lid. Take about 0·5 gm. of clear crystalline Washing Soda. Grind it to a fine powder before weighing. Heat cautiously.

Notice that copper sulphate crystals are not wet. Water of crystallization is not like water by itself. That should not surprise you. You could give many examples. The copper in copper sulphate is not like copper by itself.

The water and the copper sulphate have combined to give a Chemical Compound. It must therefore obey the Law of Constant Composition.

Experiment 24. (*a*) Weigh a crucible and add to it about 0·5 to 0·7 gm. of freshly powdered copper sulphate crystals. Find the weight of the crucible and powder.

Heat cautiously, then a *little* more strongly until the blue colour disappears. What does this show?

Weigh again and find the loss.

(*b*) With great care add water to the crucible until the powder is covered and stand on one side near or over a radiator until the next lesson. Cover with paper to keep out dust, and remove to a place of safety when school closes. (An excellent stand for crucibles is made by removing the lid and bottom from a wooden chocolate box and covering this framework with galvanized wire netting, of which the mesh is sufficiently large to hold a crucible without danger of its slipping through.)

When the blue copper sulphate appears quite dry, weigh. Compare the gain in weight in (*a*) with the loss in (*b*). Does your result agree with the Law of Constant Composition?

Lime water has been used as a test for carbon dioxide. Lime is a very familiar substance. Builders use enormous quantities in mixing mortar, which is composed of lime, sand, and water. It is also used to remove the hairs from skins in leather-making; as a fertilizer of the soil; as a disinfectant; in the manufacture of bleaching powder, &c.

As chemists we have learned that lime is calcium oxide. (Experiment 8.)

Lime is obtained from Limestone, Marble, Chalk, by heating them strongly in a kiln. The process can be imitated in the laboratory. Blackboard chalk is not **Chalk.** 'All that glitters is not gold', and all white substances are not chalk. Forty years ago schoolmasters often used real chalk, but it was unpleasantly gritty. The manufactured article used to-day still bears the name although it does not contain any chalk.

Experiment 25. Take a piece of iron wire (the wire used in electric fires answers better) and coil one end round a pencil for two or three turns. Take a piece of chalk and scrape it down until you have a splinter about an inch long and not more than a quarter of an inch thick. Weigh this carefully. Rest it in the coil, and heat it strongly for ten minutes. Allow to cool. Weigh.

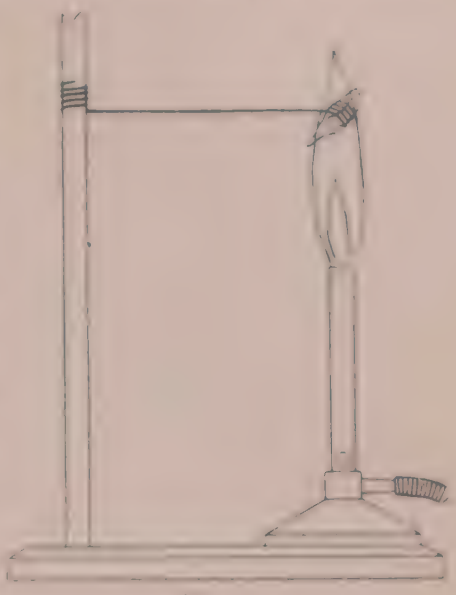


Fig. 9.

Record:

Weight of Chalk = gm.
Weight of Lime = gm.

Can you suggest a reason for the loss in weight? Is the substance lost a solid, liquid, or gas?

Manufacturers get about $11\frac{1}{5}$ cwt. of lime from one ton of chalk, i.e. 56 per cent. of the weight of chalk used. Calculate your percentage.

Chalk → Lime + a gas
56 % 44 %

Give two reasons for calling Chalk a Chemical Compound. Is Air a Chemical Compound?

Consider this very closely.

(1) Oxygen is a colourless, odourless gas.
(2) Oxygen supports combustion.
(3) Nitrogen is a colourless, odourless gas.
(4) Nitrogen does *not* support combustion.
(5) The Air contains Oxygen and Nitrogen.
(6) The Air is a colourless, odourless gas.
(7) The Air supports combustion, though not so well as Oxygen.
(8) **Air** $\left\langle \begin{matrix} \text{Oxygen} = \text{about } \frac{1}{5} \\ \text{Nitrogen} = \text{about } \frac{4}{5} \end{matrix} \right.$ but the composition varies in different localities.

You should be able to answer the question now.

There are other facts that will confirm your opinion.

(1) When Oxygen and Nitrogen are mixed together, the mixture becomes more and more like air as the proportions approach one part of Oxygen to four of Nitrogen.

(2) When air is formed by putting Oxygen and Nitrogen together, heating is not necessary, and there is no change in temperature. (Compare the making of Oxides.)

(3) Air dissolves in water, but, when the dissolved air is driven out by heating, it is found to contain Oxygen and Nitrogen in the proportion of 1 to 2, and not 1 to 4, showing that the Oxygen has dissolved more readily than the Nitrogen.

(4) When air is liquified and then allowed to boil, the Nitrogen boils away first. (Compare the boiling of water.)

From all these points we conclude that the **Air** is *not* an oxide, *not* a chemical compound, but just a mixture of gases. Chemists are very careful to distinguish between Mixtures and Chemical Compounds.

Mixed substances are readily separated.

Gold is found in the sand of certain river-beds. This sand is mixed with water in a pan, stirred thoroughly, and then the water is poured off. The water carries away a good deal of sand, but the denser gold remains at the bottom of the pan. (Density of sand 2·4 gm. per c.cm., and of gold 19·4 gm. per c.cm.) The process is repeated until only the grains of gold are left. If there is a good yield, the sand is said to 'pan out' well.

When Alcohol is made it is very dilute. It is a mixture of alcohol and water. The alcohol is separated from the water by distillation. The boiling-point of alcohol is 78° C. What is the B.-P. of water? What is a distillery?

Experiment 26 (Demonstration). Distil a mixture of methylated spirit and water from a flask fitted with a two-hole stopper carrying a thermometer. (Fig. 10.) Distil slowly and catch the distillate in test-tubes, changing the test-tube every 2 degrees. Note how the bulk of the distillate is obtained at the two boiling-points.

How could you separate (1) Mud from muddy water; (2) Salt from a mixture of salt and sand; (3) Oil from oily water; (4) Copper sulphate from a mixture of alum and copper sulphate?

Compare these separations of mixed substances with the

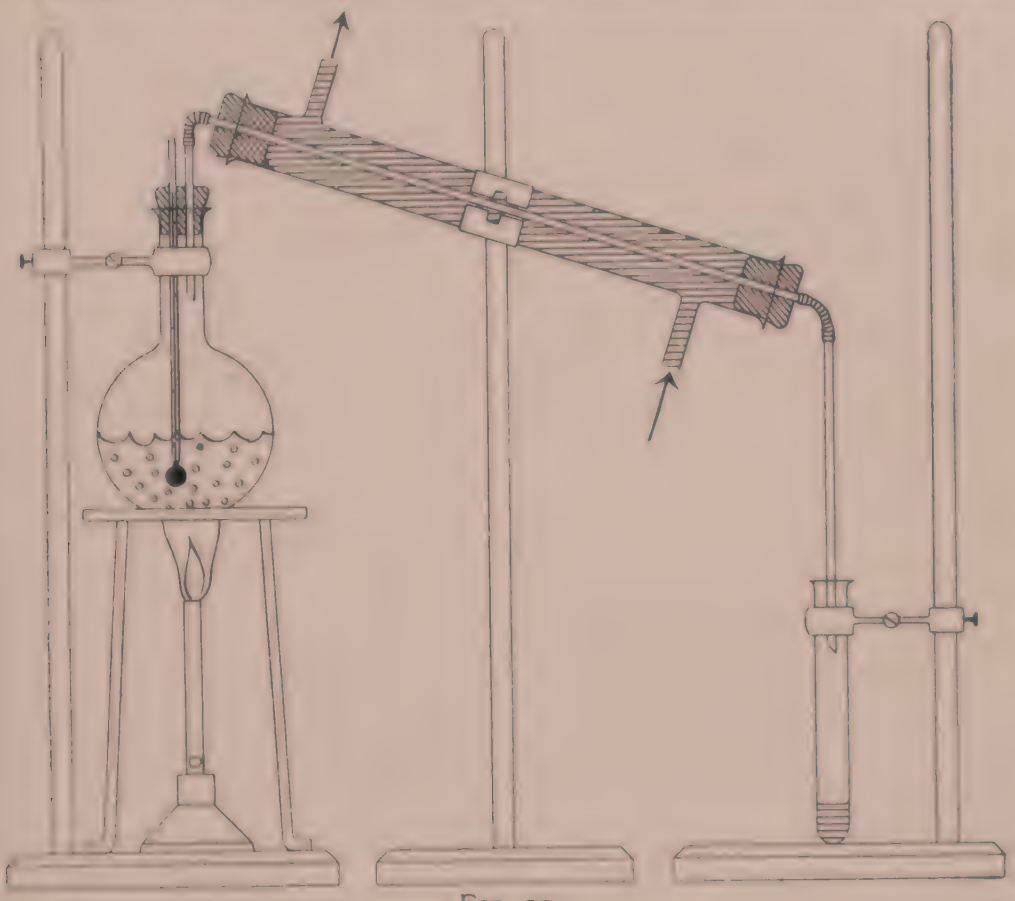


Fig. 10.

separation of the Hydrogen and Oxygen in water, or Mercury and Oxygen in mercury oxide.

Make a record of the elements and compounds you have met in Chemistry. To the lists add five you have not yet met. Ask about them before you enter them.

Elements.	*Compounds.*
Copper	Copper sulphate
Mercury	Mercury oxide
Magnesium	Magnesium oxide
&c.	&c.

Read over these quotations:

'The cinnabar of Japan has the same composition as the cinnabar of Spain. In all the world there is but one common salt. Between pole and pole compounds are identical in composition' (PROUST).

(Cinnabar is the ore from which Mercury is obtained.)

'Chemistry has discovered that the elements combine with one another in definite and unchanging ratios of quantity' (S. BROWN).

II

ELEMENTARY PHYSICS OF WATER

THE PRINCIPLE OF ARCHIMEDES

IN our everyday experience we meet many examples of articles that float on water and many that sink. More rarely we meet some that appear to lie between these two classes, e. g. a stale egg. Cooks often test eggs before using them by putting them into a bowl of water. Fresh eggs sink and bad eggs float. If the egg is not very stale it may take a long time to sink to the bottom or to rise to the surface.

Some golf balls are known as floaters and others as sinkers.

Experiment 27. Take an old golf ball that floats in water and push drawing-pins into it until it becomes a sinker. Remove drawing-pins until it cannot be described as a sinker or a floater.

Experiment 28. Take a pneumatic trough of water and drop into it a number of articles, e. g. a stone, piece of wood, a cork, a small shot. Does the weight of the article determine whether it floats or sinks? Consider the shot and the piece of wood. Does the volume determine it? Is it correct to say that large bodies float? Some large bodies do - ships; some do not— rocks.

What is it then that determines whether anything will float or sink? Can we say before putting a body into water what it will do? What information do we require before giving our opinion?

Experiment 29. Tie a piece of string to an ordinary brick, glazed, if procurable, and lower the brick into a bucket or sink *nearly* filled with water. (Why not *quite* filled?) Pull it up gradually, noting especially what happens as the brick comes out of the water. Record your experience.

Experiment 30. Take a spring balance—one reading to 12 lb.— and suspend the brick from the hook. (Fig. 11.) Note the weight of the brick.

Now take a spring balance such as is used in cooking.

This is known as a *compression* balance: the other balance is

an *extension* balance. The names indicate exactly what happens when they are used.

Allow the brick, suspended from the extension balance, to rest lightly on the pan of the compression balance, and note the readings on the two balances. (Fig. 12.)

In the diagrams a brick weighing 11 lb. is shown. Fig. 11 shows the upward pull of the extended spring carrying the whole weight of the brick.

Fig. 12 shows the upward *pull* of the extended spring carrying 8 lb. and the upward *thrust* of the compressed spring carrying 3 lb.

What is the relation of these readings to the former reading of the single extension balance giving the weight of the brick?

Allow the brick to rest more and more on the compression balance until finally the extension balance stands at 0. Note that the sum of the readings on the two balances is always equal to the weight of the brick. The weight is shared by the balances so long as they are both employed. Bear this result in mind during the next experiment.

Experiment 31 (*a*). Use the brick and extension balance but, in place of the compression balance, take a vessel of water. (Fig. 13.)

Mark the level of the water.

Lower the brick gradually into the water. Watch the balance and the water-level. Note that the extension balance behaves exactly as it did when the compression balance was used. So the weight is again being shared, this time by the balance and the water.

We can, at any point in the experiment, state the share borne by the water. What happens to this share as the brick is allowed to sink slowly into the water? What happens to the water-level? Why? Does any further change take place after the water has covered the brick, (1) in the balance-reading, (2) in the water-level?

(*b*) Find the volume of the brick. The common brick is 9 in. $\times$ $4\frac{1}{2}$ in. $\times$ 3 in. and has a volume, therefore, of 121·5 cub. in. Stand the brick on one end, and mark it off into four equal divisions, each of which will be approximately 30 cub. in. Attach the brick to the balance and lower into the water.

Since the brick receives support from the water in this

Fig. 11.

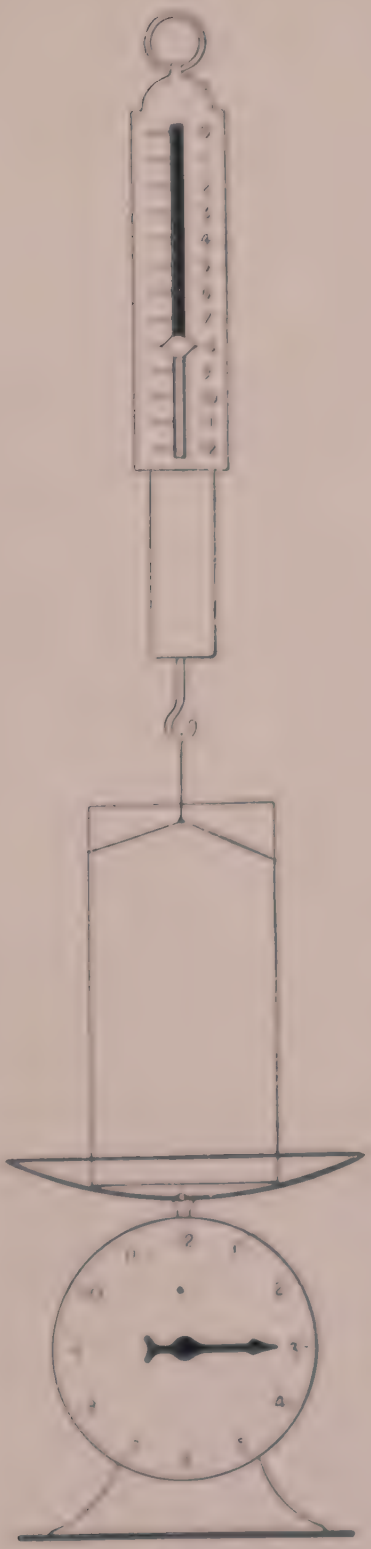


Fig. 12.

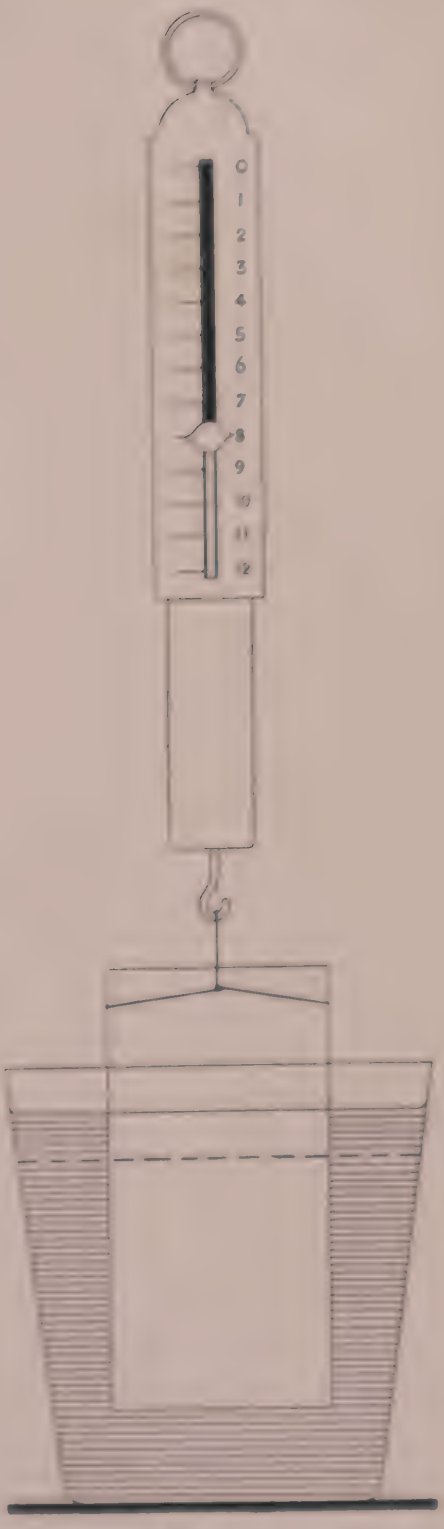


Fig. 13.

experiment as it did from the compression balance in Experiment 30, we say that the water is exerting an Upthrust. How can we tell what this Upthrust amounts to ?

Note the reading of the balance as the water reaches the marks and record :

Weight of Brick = lb.

Volume Immersed.	*Upthrust.*
30 cub. in.	lb.
60 „ „	lb.
90 „ „	lb.
The Whole.	lb.

Water exerts an Upthrust on bodies immersed. (In Science the word 'body' is used to signify a piece of matter.)

This Upthrust increases as the volume of the body below the surface of the water increases.

Experiment 32. Repeat the experiment with a block of wood—if possible, the same shape and size as the brick. Mark it off into four equal sections, and note the share borne by the water as each section is submerged.

Record.

Weight of wooden block = lb.

Volume Immersed.	*Upthrust.*
30 cub. in.	—
—	—
—	—

How does this experiment differ from the previous one ?

For **Sinking Bodies** the upthrust, when the body is totally submerged, is less than the weight of the body, i. e. the force due to the water, acting upwards, is less than the force of gravity, acting downwards, and so they sink.

Notice what happens when you try to push the wooden block below the surface with your finger, and write out a statement about floating bodies, using the preceding statement about sinking bodies as a pattern.

In the light of these experiments think over the following:

A large stone is easily moved *under* water.

A diver quickly comes to the surface.

A fisherman uses a landing-net for his big fish.

A fisherman often speaks of losing his heaviest fish as it is being pulled out of the water.

A swimmer is just able to keep his head above water.

We noticed that there was evidently some relation between the volume of the brick under water and the upthrust. See record, Experiment 31 (*b*).

Calculation.

The volume of the brick is 121·5 cub. in. (Exp. 31 (*b*).)

1 cub. ft. water weighs 1,000 oz.

or 1,728 cub. in. of water weigh 1,000 oz.

$\therefore$ 121·5 cub. in. of water weigh $\frac{1{,}000}{1{,}728} \times 121{\cdot}5 = 70$ oz.

The weight of water having the same volume as the brick is 70 oz. or $4\frac{1}{2}$ lb. (nearly).

Turn to the result of Experiment 31 (*b*), and compare the weight of the water having the same volume as the brick with the upthrust due to the water when the whole brick was submerged. What do you notice? State it.

WARNING.

In dealing with the results of experiments involving measurement it is necessary to take into account the degree of accuracy of the instruments used. In the experiment just discussed the brick was measured roughly. Its volume is *about* 121 5 cub. in. The weight was found on a common spring balance. Notice what each division on the scale stands for. It would be quite unscientific to pretend that

the results are very accurate, and use them in this way, for example:

$$
\begin{array}{r}
1728)\,121500(70{\cdot}3125\text{ oz.} \\
12096 \\
\hline
5400 \\
5184 \\
\hline
2160 \\
1728 \\
\hline
4320 \\
3456 \\
\hline
8640 \\
8640 \\
\hline
\end{array}
$$

Ans. The weight of water having the same volume as the brick = 70·3125 oz.

Before using numerical results always consider how far they may be trusted. It is clear we cannot trust '0·3125 oz.' in the example given.

In studying the air, we spoke of **Matter** occupying space or taking up room. When a bottle is filled with water the air has to find a way out before the water can get in. We say that the water displaces—takes the place of—the air. In the lowering of the brick into water, the level of the water rose, i. e. the brick displaced water, and the volume of the water displaced was equal to the volume of the brick below the surface of the water.

So we may write:

The weight borne by the water
or
The upthrust of the water on the brick
} depends upon the volume of water displaced by the brick.

The volume of displaced water was found to be 121·5 cub. in., and its weight $4\frac{1}{2}$ lb. The balance showed that the upthrust was $4\frac{1}{2}$ lb.

$\therefore$ **The upthrust on a body immersed in water** $=$ **Weight of water displaced by that body.**

This statement is usually known as the

Principle of Archimedes.

Archimedes (287–212 B. C.) probably received part of his education under Euclid at Alexandria. He discovered the principle of the lever (Bk. I, p. 21), and declared that if he had a lever long enough he could move the world. The story of his discovery of the principle which bears his name is an interesting one.

Hiero, the King, engaged a goldsmith to make a crown and provided the gold for the purpose. When it was delivered the King suspected the man of keeping some of the gold and substituting silver for it. The weight, however, was correct. The King set Archimedes the task of discovering whether the goldsmith had acted dishonestly.

It is said that one day Archimedes stepped into his bath, which was full to the brim, and, noticing the overflow, shouted 'Eureka' ('I have found it'). He saw his way to solve the King's problem.

He obtained two balls, one of silver, and one of gold, each the same weight as the crown. He then compared their volumes by marking the level of water in a vessel and, putting each in turn into the water, he noted the new levels. He saw at once that the volume of the gold was less than the volume of either the crown or the silver, and that the volume of the silver was greatest. He concluded that the crown was made of silver and gold, and not of pure gold. He even went a step farther, and had a number of balls made of an alloy of gold and silver, mixed in different proportions, but all of the same weight as the crown, until he found a ball that caused the water in the basin to rise to the same level as it did when the crown was immersed. Archimedes could thus tell the King the amount of silver substituted for gold in the crown.

The Principle of Archimedes

Archimedes was so absorbed in the work that he continued to study the behaviour of bodies in water until he was able to show the truth of what has been named **The Principle of Archimedes.**

While working out a geometrical problem on the sand Archimedes was killed at Syracuse by a Roman soldier.

The experiments on Upthrust will now be repeated with scientific apparatus.

To carry out an experiment with a body totally immersed in water demands special care. A beaker of water is placed on a small table which stands astride the left-hand pan of the balance. In using this make quite sure that neither the beaker nor the table touches the balance.

Six or eight pieces of glass rod $1\frac{1}{2}$ in. long, tied up into a bundle, make a very convenient body for weighing. Such a body, used in experiments on the Principle of Archimedes, is often called *the sinker.*

Experiment 33. Place the small table astride the balance pan. Suspend the sinker, i. e. the bundle of glass rods, from the balance-hook by means of fine cotton thread (No. 60), and cut off any loose ends. (To obtain very accurate results some scientists use unspun silk or even a hair.) The sinker should be within a $\frac{1}{4}$ in. of the small table. (Fig. 14.)

Add sand or lead shot to the right-hand pan until the sinker is just balanced. A few small pieces of paper may be used at the end to make the balance exact. Pull the sinker to one side, stand a beaker of water on the table, and allow the sinker to hang freely in the water. Raise the balance pans, and note the effect of the water on the sinker. The pans do not swing because the sinker is pushed up by the water; it is experiencing an upthrust. To find how great this upthrust is, add weights to the L.-H.P. (left-hand pan) until the balance again swings. The added weights have thus cancelled the upthrust and are equal to it.

(A) ∴ Upthrust due to water = gm.

Now find the volume of the sinker.

Pour some water into a burette. Take the reading very care-

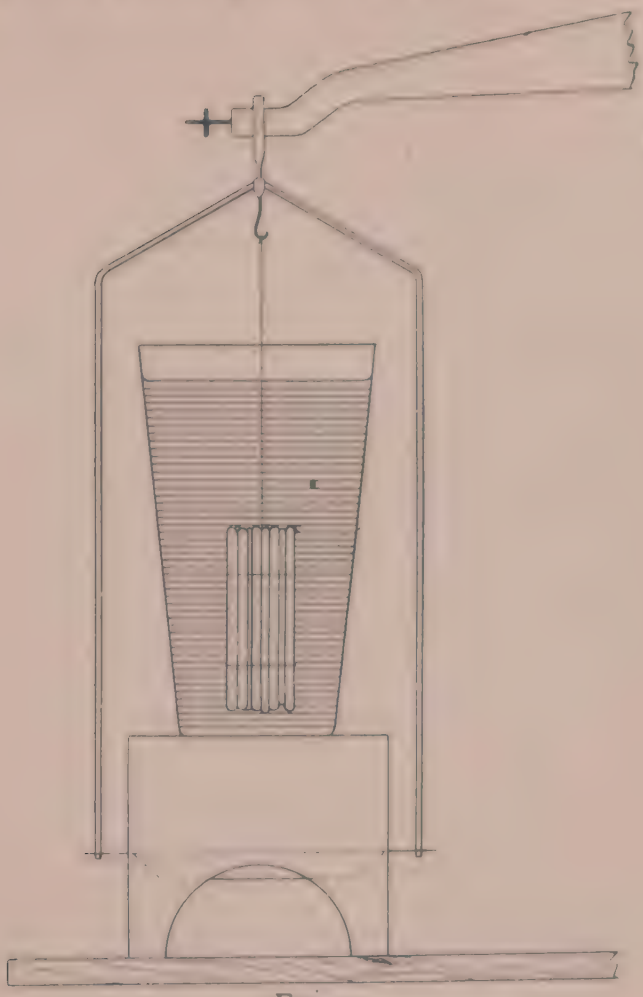


FIG. 14.

fully. Drop in the pieces of glass rod, singly, and take the reading again.

Record.

Burette reading = c.cm.

Burette reading after adding sinker − c.cm.

Volume of water displaced by sinker and ∴ volume of sinker = c.cm.

But 1 c.cm. water weighs 1 gm.

(B) ∴ Weight of water displaced by sinker = gm.

Look at the two results (A) and (B), and state your conclusion.

Keep the sinker for future use.

The method usually adopted is given in the next experiment.

Experiment 34. Dry the pieces of glass rod and tie them up into a bundle again. Suspend them from the balance hook as before.

Instead of using sand or lead shot, find the weight of the sinker in the ordinary way.

Weight of sinker = gm.

Remove the weights, allow the sinker to hang in water, and again add weights to the R.-H.P. as in ordinary weighing. The weights required are less than the weight of the sinker. (Why?) By how much? Subtract to find this.

Weight of sinker	=	gm.
Weights used when sinker is immersed in water	=	gm.
Difference, i. e. Upthrust due to water	=	gm.

Compare this result with (A) of the last experiment.

SPECIFIC GRAVITY

We have already used the word Density, and defined it as the mass in grammes of 1 c.cm. of a substance $\left[\frac{M}{V}\right]$. The upthrust of the water on a sinker has been found to be equal to the weight of the water displaced by the sinker, e.g. in an experiment this record was obtained:

Weight of glass sinker = 18·4 gm.
Weight of water displaced by the glass sinker = 8 gm.

and so the

Volume of water displaced by the glass sinker = 8 c.cm.
8 c.cm. glass weigh = 18·4 gm.
and ∴ 1 c.cm. glass weighs 18·4 ÷ 8 = 2·3 gm.

Density of glass is 2·3 gm. per c.cm.

The water displaced by the sinker has the same volume as the sinker, so that the two weights entered in the record refer to equal volumes, viz. 8 c.cm., of different substances —glass and water. (Fig. 15.)

The glass and the water have both the same volume, and the glass weighs $\frac{18{\cdot}4}{8}$ or 2·3 times as much as the water. Its Density is 2·3 times as great as the density of water.

Water is a very convenient substance: it is plentiful; it is colourless; it is odourless; it is transparent; it is easily obtained pure; it does not cost anything. Scientists have agreed to use *water* as one of the standards to measure other things by. You will recall that the gramme weight depends upon water. When comparing the densities of substances reference is made to water as the standard; they are considered with respect to the **Mass** (Bk. I, p. 21) of the same volume of water, and the number expressing this is termed **Specific Gravity**. The **Specific Gravity** of **Water** is taken as the **Unit**, i. e. 1.

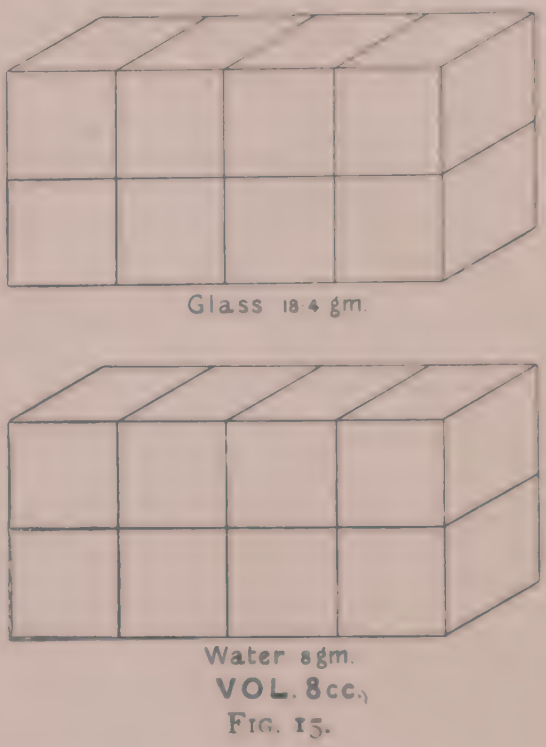


FIG. 15.

The mass of 1 c.cm. water is 1 gm. The Specific Gravity of Glass is 2·3, for 1 c.cm. glass has a mass of 2·3 gm.

This is written:

s.g. Water	1 (The Standard.)
s.g. Glass	2·3

Turning to the results we found under Exp. 10, Bk. I, we may write:

s.g. Methylated spirits . .	0·85
s.g. Turpentine	0·87
s.g. Paraffin	0·83
s.g. Sea-Water	1·025
s.g. Mercury	13·6

Remember *specific gravity* is just a number—it does *not* stand for grammes.

A gallon jar holds 10 lb. of water.

If a gallon jar is filled with mercury, the mercury will weigh 13·6 times as much, i. e. 10 lb. × 13·6 or 136 lb. A 2-gallon petrol tin would hold 20 lb. of water. It therefore holds 20 lb. × 0·8 or 16 lb. of petrol, because the s.g. of petrol is 0·8.

The method of Experiment 34 enables us to find the s.g. of any solid substance that is insoluble in water and sinks, by using the Substance as a sinker.

Experiment 35. Find the s g. of one of the following substances: Lead, Copper, Iron, Coal, Marble, Brass. Each should be found by two students independently, to act as a check, and the results tabulated.

Archimedes could have settled the point about the crown more readily if he could have found the s.g. of gold, of silver, and of the alloy. Gold is 19·3 times as heavy as water of the same volume. Silver is 10·5 times as heavy as water of the same volume.

or s.g. Gold . . . 19·3
s.g. Silver . . . 10·5

If the dishonest goldsmith had taken one bar of gold and one bar of silver, having the same volume, and melted them together to form the alloy from which he made the crown, then the specific gravity of the alloy would have been 14·9,

i. e. $\frac{19{\cdot}3 + 10{\cdot}5}{2} = 14{\cdot}9$.

Articles made of standard silver, usually spoken of as solid silver, and British Coinage up to 1920, have 92·5 per cent. silver and 7·5 per cent. copper, or 925 parts of silver in 1,000.

In 1920 the percentage of silver in coinage was reduced to 50 per cent., or only 500 parts of silver in 1,000.

Experiment 36. Obtain four two-shilling pieces or half-crowns dated earlier than 1920 and another set dated later than

1920. Tie each set of coins in a bundle, and find the s.g. of the alloys of which they are made by the method of Experiment 34.

Record.

Silver Coins minted before 1920.

(A) Weight of standard silver = gm.
Weights used when coins are immersed in water = gm.
Difference, i.e. Upthrust due to water, and ∴ (B) Weight of water displaced = gm.
(A) Weight of coinage silver = gm.
(B) Weight of same volume of water = gm.
s.g. of coinage silver minted before 1920 =

From another member of the class you should get the result for the other set of coins.

s.g. of coinage silver minted since 1920 =

N.B. Specific Gravity is concerned with *matter*, the material or stuff of which a thing is made, and not with the article fashioned from the matter, e. g. we may speak of the s.g. of brass, but not of the s.g. of a candlestick ; of standard silver, but not of a florin.

Specific Gravity is the number that expresses how many times the mass of a substance is greater or less than the mass of an equal volume of the standard substance, water.

$$\text{or s.g.} = \frac{\text{mass of substance}}{\text{mass of equal volume of water.}}$$

THE PRINCIPLE OF ARCHIMEDES APPLIED TO OTHER LIQUIDS

Consider the glass sinker. You know its volume, the weight of water it displaces, and its weight. If you allow it to hang in another liquid, e. g. methylated spirit, will the upthrust be equal to the weight of methylated spirit displaced, just as when it hangs in water the upthrust is equal to the weight of water displaced? This can be decided by experiment.

Several of the common liquids should be used by the class, and a class record written on the blackboard. The volume of the sinker is known (Exp. 33), and therefore the volume of the liquid it displaces. The weight of the liquid displaced can be found if its Density is known.

Experiment 37. Calculate and enter in your record (A) the weight of the liquid displaced by your sinker. Then, by the method of Experiment 34, find (B) the upthrust of the liquid on the sinker.

Thus : e.g. **Turpentine.**

Volume of sinker and volume of turpentine displaced	=	8 c.cm.
Density of turpentine	=	0·87 gm. per c.cm.
(A) Weight of turpentine displaced (8 × 0·87)	=	6·96 gm.
(B) Upthrust of turpentine	=	

Class result.

Liquid.	*Weight of Liquid displaced.*	*Upthrust.*
Turpentine	gm.	gm.
Methylated Spirit	gm.	gm.
Paraffin	gm.	gm.
Sea Water	gm.	gm.

What do you conclude from your results?

The Principle of Archimedes is thus true for liquids generally. It is also true for gases.

The Hydrogen balloon rose to the ceiling because the upthrust of the air in which it was immersed was greater than the weight of the balloon.

e.g. Weight of 4,000 c.cm. Hydrogen	= 0·36 gm.
Weight of Rubber Balloon	= 3·00 gm.
Total Wt.	= 3·36 gm.

1,000 c.cm. air weigh 1·24 gm.

∴ 4,000 c.cm. air displaced by the balloon weigh 4·96 gm., and the upthrust on the balloon = 4·96 gm., which is greater than the weight of the balloon.

This can all be summed up as follows:

The Principle of Archimedes states that when a body is immersed in a fluid the upthrust is equal to the weight of fluid displaced. (Fluid, Lat. *fluere*, to flow, i. e. liquids and gases.)

FLOTATION

So far we have worked with a 'sinker'. We shall now experiment with a floating body.

Experiment 38. (A) Take a test-tube and load it with lead shot so that it floats upright in water. Do not put in much more shot than is required. Float the loaded test-tube in water in a measuring jar. (Fig. 16.)

Note the reading of the water-level, before and after putting in the test-tube, to find the volume of water displaced by it. Lift out carefully, wipe dry, and weigh.

Record.

1st reading on measuring jar = c.cm.
2nd reading on measuring jar = c.cm.

∴ Volume of water displaced by floating body = c.cm.
∴ Weight of water displaced by floating body = gm.
Weight of floating body = gm.

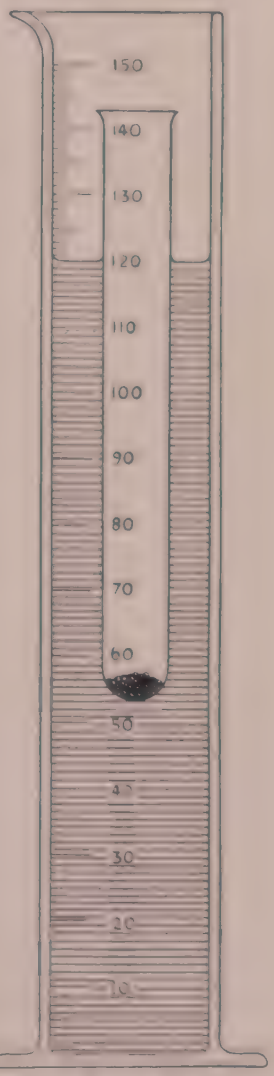


FIG. 10.

Float the loaded test-tube again. Tie a 5 gm. weight (or two 2 gm. wts. and a 1 gm. wt.) to a thread, and lower it into the test-tube, allowing it to rest on the lead shot. What

difference does this make to the volume and weight of water displaced?

Read the warning on page 37.

What conclusion do you draw from this experiment?

Experiment 39. Put salt solution or sea-water in the measuring jar. The density of this has already been found (Bk. I, Exp. 10). Float the same loaded test-tube in the salt water, and record as before.

Record.

1st reading of measuring jar	=	c.cm.
2nd reading of measuring jar	=	c.cm.
Volume of salt solution displaced	=	c.cm.
but 1 c.cm. salt solution weighs gm.		
∴ Weight of salt solution displaced	=	gm.
Weight of floating body	=	gm.

Look carefully at the results of the last two experiments. Note the difference in the volume of liquid displaced.

Into which does the floating body sink more deeply, the water or the salt water? Why?

Consider the following questions: What happens as the cargo is put on board ship? What happens when a ship loaded at a river port sails out to sea? What happens when a ship loaded at a seaport sails up a river?

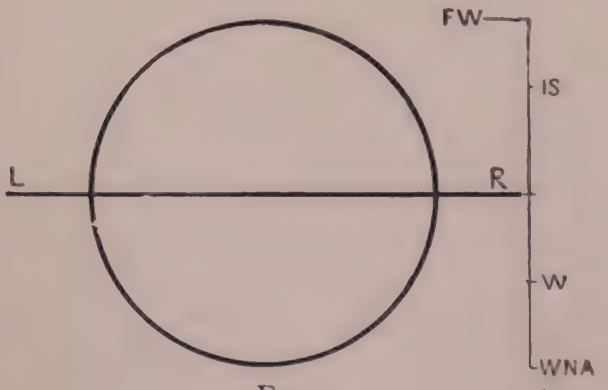


Fig. 17.

The Plimsoll Mark, or Plimsoll's Eye, as sailors call it, is a circle 12 in. in diameter with a horizontal line 18 in. in length drawn through its centre. This is the legal load-line for the ship, and when loaded the ship must not sink below that line. The Plimsoll Mark is a result of the untiring effort of Samuel Plimsoll on behalf of sailors; and to him they owe

the Merchant Shipping Act, 1876, which among other rules compels every British merchant ship over eighty tons to show the mark on each side. (Fig. 17.)

In the figure L R stand for Lloyd's Register. 'The upper edge of the horizontal line passing through the centre of the disk shall always indicate the summer load-line in salt water.' To the right of the 'mark', ships often show other levels marked F W, i.e. fresh water, I S, i.e. Indian Summer, W, i.e. Winter, and W N A, i.e. Winter North Atlantic.

Why are two of these above and two below the chief line?

HYDROMETERS

The method of the floating test-tube is sometimes used to find the s.g. of a liquid. An instrument based on the same principle is known as a Hydrometer, Lactometer (Lat. *lactis*, milk), Alcoholometer, according to its use. (Fig. 18.) Note the weighted end, the enlarged portion of the stem, and the narrow portion containing the scale. Think out the reason for each of these.

Fig 18.

Experiment 40. Take a strip of paper with millimetre ruling and number the centimetres plainly. Round off the end so that it fits closely to the bottom of a test-tube. Load the test-tube to make it float upright in water. (Fig. 19.)

Read the number on the scale at the water-level.

Record.

Lift the test-tube out of the water, wipe, and put it into alcohol, turpentine, paraffin, brine, or other liquid. Read the scale and record.

Since the weight of the floating body has not been changed, the weight of liquid displaced will equal the weight of water displaced. (Why?) Therefore

the water equal in volume to the part of the test-tube submerged in the water has the same weight as the liquid equal in volume to the part of the test-tube submerged in the liquid.

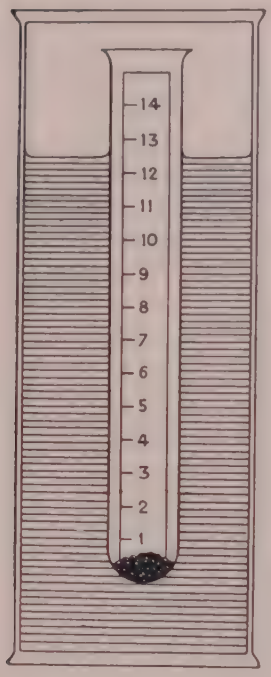


Fig. 19.

In an experiment the reading for water was 12 and for turpentine 15.

The weight of water displaced by 12 cm. of the test-tube was the same as the weight of turpentine displaced by 15 cm. of the test-tube. It is clear that turpentine is less dense, or has a smaller specific gravity than water. In fact its s.g. is $\frac{12}{15}$ of the s.g. of water, or 12 times the s.g. water = 15 times the s.g. turpentine.

The s.g. water is 1.

$\therefore$ 12 × 1 = 15 × s.g. turpentine.

$\therefore$ $\frac{12}{15}$ or 0·8 = s.g. turpentine.

It is because this method is carried through so quickly that it is very commonly used in industry, as, for example, in finding the s.g. of whisky and other spirituous liquids, of milk, of oils, of accumulator acid, &c.

The Principle of Archimedes also provides us with another and very reliable method for finding the Specific Gravity of a liquid.

A sinker is weighed (1) in air, and (2) immersed in water, and the weight of a volume of water equal to the volume of the sinker is thus found.

The sinker is then weighed (3) in the liquid, and the weight of an equal volume of the liquid is found.

Experiment 41. From the information in the preceding paragraphs, use the Principle of Archimedes to find the Specific Gravity of a liquid.

Why does wood float? Why does iron sink?

A cocoa tin is closed and is found to float on water. It is crushed flat. Will it float or sink? State the reason for your answer very fully.

How is it possible for an iron ship to float?

Some of you may like to prove the Principle of Archimedes by reasoning it out.

Imagine one cubic centimetre of water in the middle of the water in a beaker. The water is quite still in the beaker, so that our cubic centimetre is neither rising to the surface nor sinking to the bottom. It is just supported and no more, and therefore the upthrust on it is equal to its weight. If we put in place of our cubic centimetre of water a cubic centimetre of glass, the upthrust will be just the same, that is, the upthrust on the glass will equal the upthrust on the water that was there before the glass was put in. (Fig. 20.)

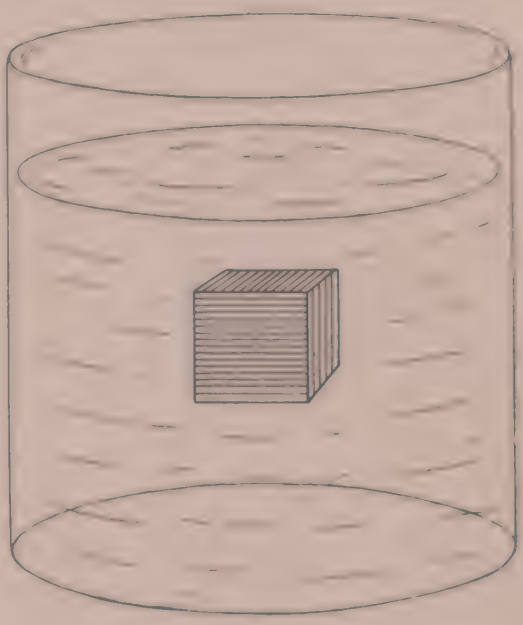


Fig. 20.

A liquid, like a gas, exerts pressure equally in all directions. The pressure at any point on the liquid depends upon the depth. The pressure on the sides of the cube are equal and opposite and so neutralize one another. Since the bottom of the cube is 1 cm. lower than the top, the upward pressure on it is greater than the downward pressure on the top by the weight of 1 c.cm. of water: i.e. the upthrust on a body whose volume is 1 c.cm. equals the weight of 1 c.cm. of water. (Fig. 20.)

It is quite likely that Archimedes used reasoning such as this.

HEAT

Transmission of Heat. I.

In measuring liquids it is important to guard against loss by leakage—a quart cup with a hole in it would not make a satisfactory measuring vessel; it must be water-tight. To test a measuring vessel and make quite sure that it does not leak is quite a simple matter.

Pneumatic tyres, footballs, tennis-balls must be air-tight.

For measuring **Heat** we require *heat-tight* vessels, and before we can stop heat from leaking we must know how heat leaks.

The purpose of our clothing is to prevent heat leakage. In winter we increase our efforts to conserve the heat of our bodies.

What is an eiderdown quilt for? Why do we use blankets? Why do motorists put a rug over the bonnet of a car? What is the radiator for? What is a Thermos flask for? Why do *some* people pour hot tea into the saucer? What is a hay-box?

Think over these examples. In some, heat leakage is being hindered, in others, helped.

We have spoken of temperature as heat-level (Bk. I, p. 39) and naturally we should expect things in contact with one another to come to the same-heat level or temperature. If you hold an iron rail, your hand soon feels cold. Heat is leaking from your hand to the iron. When a poker is left with one end in the fire for some time, the other end of the poker also becomes hot. Heat passes along the metal from the hotter to the cooler portion.

Experiment 42. Hold the edge of a florin at the side of a bunsen flame and notice how quickly the heat travels across the coin. Try a copper coin and pieces of any other metals available.

Hold a small piece of burning paper and notice how near the flame gets to the fingers before you are compelled to drop it.

Try a piece of wood similarly. Look very carefully and say exactly what causes you to drop the paper and wood.

Experiment 43. Take a little melted paraffin wax and drop spots of it at intervals of half an inch along the rod of a retort stand. Before the wax sets, put the head of a tin-tack into each drop and hold the tin-tack at right angles to the rod until it is firm. Fit up the apparatus as in Fig. 21. Make a hole in a piece of

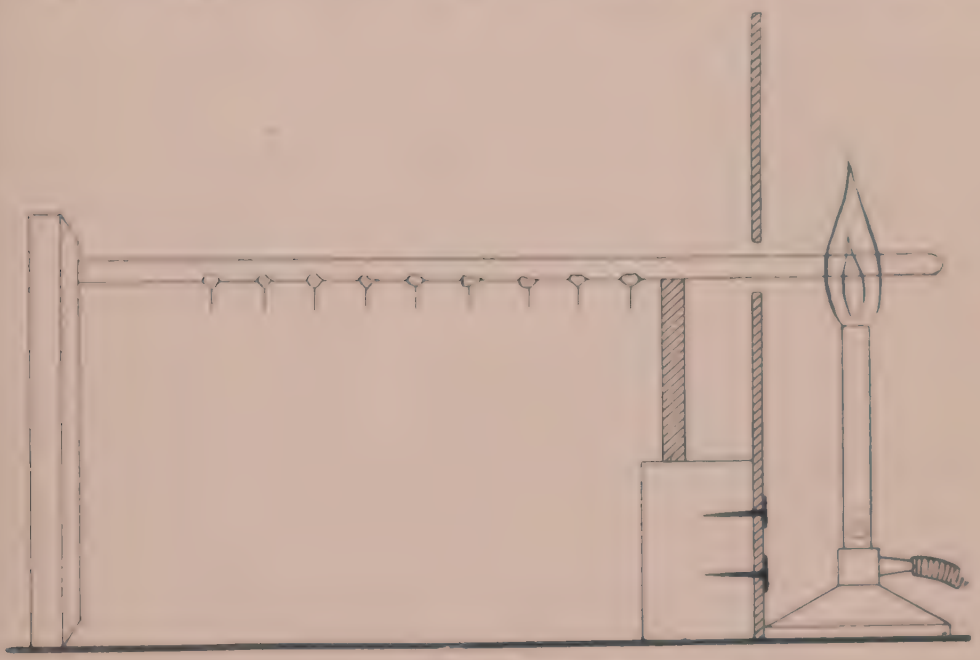


Fig. 21.

cardboard, and pass the end of the rod through it to shield the apparatus from the Bunsen flame. Now heat the end of the rod, and note the result.

This experiment shows that the heat is passing along the rod from the hotter to the cooler end and heating the rod, bit by bit, on its way. The process by which heat is passed along from one particle of a body to the neighbouring cooler particles, in *contact* with it, is called **Conduction.** (Lat. *con*, together; *ducere*, to lead.)

Substances which allow heat to pass along readily are called Good Conductors, e.g. silver, copper, and metals generally.

Substances which appear to offer resistance to the passage

of heat are called Bad Conductors, e. g. wood, paper, glass, water, air.

The Miners' Safety Lamp. This Safety Lamp depends upon two facts : (1) that combustible substances will not burn unless they are hot enough ; (2) that metals are good conductors of heat.

Experiment 44. Regulate the air holes of the Bunsen burner so that the flame is only moderately hot. It should be almost

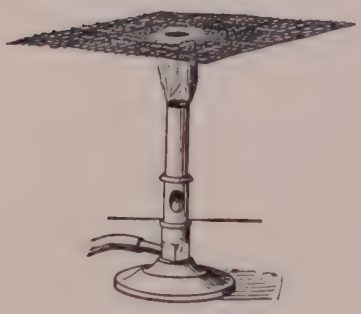


Fig. 22.

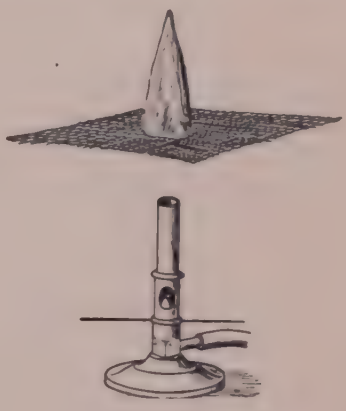


Fig. 23.

a luminous flame. Bring down a piece of wire gauze on the flame and note that the flame does not pass through to the upper side.

Turn off the gas and relight with the flame above the gauze. Again the flame does not pass through. (Fig. 23.)

The wire gauze is composed of many iron wires, and iron, a metal, is a good conductor of heat. Each wire is a path along which the heat may travel. The heat is conducted away so quickly that the combustible gas on the other side of the gauze is not heated to a temperature high enough to ignite it.

Experiments of this nature set Sir Humphry Davy, in 1815, thinking of a 'flame sieve'. A sieve sifts or separates ; some things get through, others do not. Explosions were all too common in coal-mines, and he saw that with a flame sieve, i. e. a wire gauze, a flame might burn safely although it was

only separated from inflammable gas by the gauze. Fresh air could pass to a flame surrounded by wire gauze and the products of combustion could get away. His Lamp (Fig 24) was an oil-lamp surrounded by a cylinder of wire gauze. Its light was feeble but safe. Though it has been much improved since his day, the miners' lamp still has its flame protected by a 'flame sieve'.

Examples of conductors and bad conductors are quite common.

A wooden spoon is never too hot to hold—a silver spoon handle gets very hot.

Tea-pots and kettles often have the handle partly made of wood.

Hot-water taps are sometimes fitted with a wooden cross-piece at the top.

The heat travels quickly through a frying-pan.

Iron-holders and kettle-holders are used to protect the hand.

In Conduction the heat is led along the material, the body itself remaining still.

Fig. 24.

Transmission of Heat. II.

We have mentioned that water is in the class of bad conductors.

Experiment 45. Take a small piece of ice and wind wire round it so that it will sink in water. Put the prepared ice into a test-tube of water. Turn the Bunsen flame down and hold the topmost inch of water at the flame. When the water boils, look at the piece of ice at the foot. When it is heated from below, instead of from above as in this experiment, the water soon boils, as a whole.

In conduction heat travels in all directions. Water is most quickly heated from below. The next experiment will explain this.

Experiment 46. Place a beaker of water on a tripod stand, without a gauze. Drop in one crystal of potassium permanganate, to one side if possible, and bring a *small* flame immediately below the crystal but not touching the glass. Watch the track of the coloured water, and make a sketch of it.

A liquid, then, does not remain still when it is heated. It begins to stir itself and carry the heated water upwards, while cooler water passes down to the source of heat. This is a very effective way of distributing the heat and is known as **Convection.** (Lat. *con*, together; *veho*, I drive.) The water acts as a *vehicle* conveying heat throughout the whole quantity. Convection is only possible in fluids. It is impossible in solids. Why?

Can we explain convection? Consider the experiment. The water at the foot near the crystal of potassium permanganate was heated first. Water, on heating, expands; its density becomes less, so that cooler and denser water is resting on warmer and less dense water.

The denser water sinks and becomes heated, while the less dense water rises. Thus a continual stirring is set up and all the water becomes heated.

Houses and Public Buildings are frequently heated by Convection Currents. (Fig. 25.) The tank in the roof allows for the expansion of the water, and the pipe shown serves as a way of escape for steam.

Ocean Currents are convection currents on a large scale. Look at a map showing Ocean Currents. Where is the water heated?

Convection Currents in gases are readily illustrated.

Heated air from the fire passes *up* the chimney.

Cigarette smoke curls *upwards.*

The air near the ceiling is always warmer than the air near the floor.

Air is a very bad conductor of heat, but convection currents are quickly set in motion. If then we wish to use the air to prevent heat leakage we must do something to

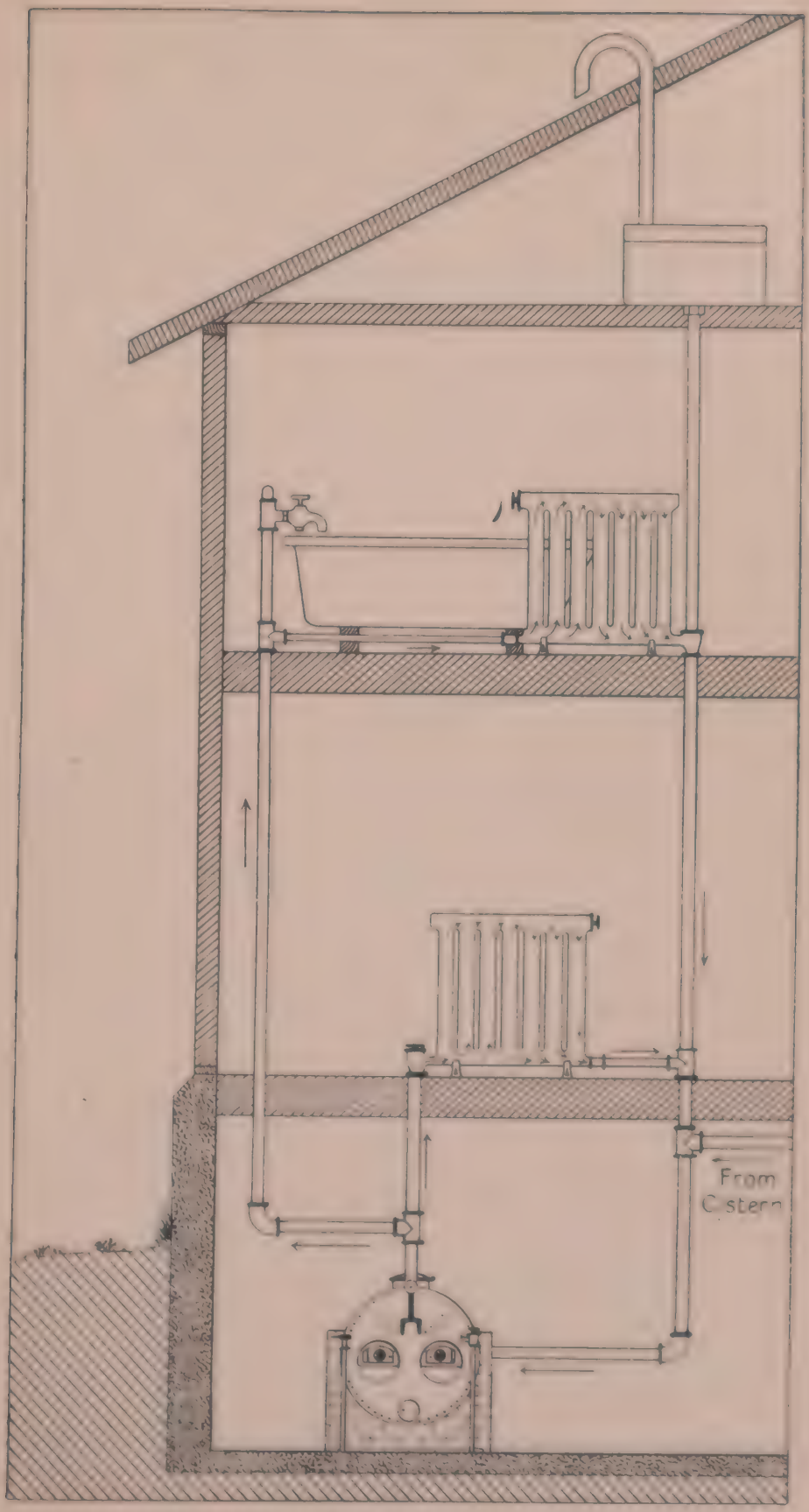


Fig. 25.

entangle it and prevent it from circulating. We must hinder Convection.

A tight kid glove is *cold* in cold weather because it fits too closely to the hand. A larger glove, woollen lined, entangles a good deal of air, which, acting as a bad conductor of heat, keeps the warmth of the hand from escaping.

Many examples will occur to you: the wool of the sheep, hair and fur of animals, our own clothing, blankets, feather pillows. Even snow, or rather the air entangled by the snow, prevents the loss of heat from the ground during severe frost.

Polished surfaces *feel* colder than rough ones because, when we place the hand on a rough surface, all the tiny hollows are full of imprisoned air.

A kettle-holder, a tea-cosy, and a cushion are largely made up of air so imprisoned that convection currents are prevented.

Transmission of Heat. III.

How is it that we feel warmed when sitting before a fire?

Air is a very bad conductor of heat so *conduction* does not explain it.

Heated air rises and cold air takes its place. The heated air will not reach us so *convection* does not explain it. In fact, convection *does* explain why we so often complain of a draught at our backs when we are close to the fire. If the fire is too fierce we hold a screen between it and our faces to stop the heat.

The sun's rays do not reach us when we stand in the shade; it is darker and cooler than in the sunshine. When we shut out of sight the source from which the heat proceeds we also cut off the heat, because heat travels in the same way as light. By placing an object in front of a lighted lamp, a shadow is cast, and in exactly the same way we can produce a heat shadow. Heat, then, travels in rays as light does, passing outwards in all directions from the source. This is called **Radiation.** (Lat. *radius*, spoke of a wheel.) It is quite unlike Conduction and Convection.

Experiment 47. Take a magnifying glass, sometimes known as a burning-glass, and hold it facing the sun. Bring a piece of paper gradually nearer the lens and parallel to it. Notice how the circle of light becomes smaller and brighter. The light rays are being brought to a focus.

When the light spot is as small as you can get it the paper begins to scorch, showing that heat rays, too, are brought to a focus.

Heat rays are therefore concentrated as well as light rays. Just as all the light rays falling on the glass are brought to a point so are all the heat rays. This very bright and very hot spot is called the focus. (Lat. *focus*, a hearth.)

Priestley used a burning-glass to heat the mercury oxide from which he obtained Oxygen.

As a barometer is taken up a hill the mercury falls, because the air only extends to a comparatively short distance from the earth. (Bk. I, p. 85.) The space between the earth and the sun contains no matter at all, it is a vacuum, yet the light and heat of the sun travel to the earth. It is interesting to note that they both travel at the same rate. When there is an eclipse of the sun, it becomes darker and colder at the same time. Even a cloud passing between us and the sun causes the air around us to feel cool.

Experiment 48. Take a piece of old wire gauze and fold it over and over until you have a small packet about 1 in. square by $\frac{1}{4}$ in. Twist one end of a wire round the packet to make a handle, and support it from the rod of a retort stand. Place the stand so that the gauze is heated to *bright* redness in the Bunsen flame.

Take two stiff covers from old exercise books and make a folding screen to shut in the Bunsen and gauze on three sides. The screen should now resemble a clothes-horse placed before a fire.

Put your hand in the enclosure.

Place a thermometer in the enclosure and watch the mercury.

A tiny feather or smoking paper will show that convection currents are carrying the heated air upwards, so that even if air

were a good conductor of heat it would not be possible for it to conduct heat to the screen.

Heat is transmitted from the gauze to the screen by **Radiation.**

Experiment 49. Take a metal box (a 2-oz. tobacco tin, cylindrical, answers well) and make a hole in the side large enough to admit a thermometer with a short piece of rubber tubing round the stem to hold it firmly in the hole. Put another piece of rubber tubing at the top of the stem to protect it when held in an iron clamp. Grip this rubber in the clamp of a retort stand so that the thermometer hangs vertically with its bulb inside the metal box.

Place the apparatus thus prepared inside the screen of Experiment 48. The bright bottom of the tobacco tin must be about 6 in. away from the red-hot gauze. (Fig. 26.)

When the thermometer remains steady, record the temperature.

Remove the thermometer and tin without disturbing the rest of the apparatus. Set fire to a piece of camphor or naphthaline (size of a pea) and hold the bright tin over it until it is covered with soot. Put the apparatus back inside the screen exactly as it was before, and when the thermometer remains steady, record the temperature.

Record.

Temperature when bright tin faces heated gauze = °C.
Temperature when dull black tin faces heated gauze = °C.

What conclusion do you draw from your results?

Schoolboys are all familiar with the small mirror which, when held in the sunshine, reflects the light to the walls or ceiling of the room. The brighter the object the more light is reflected. Reflectors used in cycle and car lamps are either glass mirrors or highly polished metal. When a cycle lamp mirror becomes covered with soot it throws very much less light on the road.

Radiant Heat is like Light. It can be reflected from bright surfaces or absorbed by dull ones.

White clothes are much worn in tropical countries. In

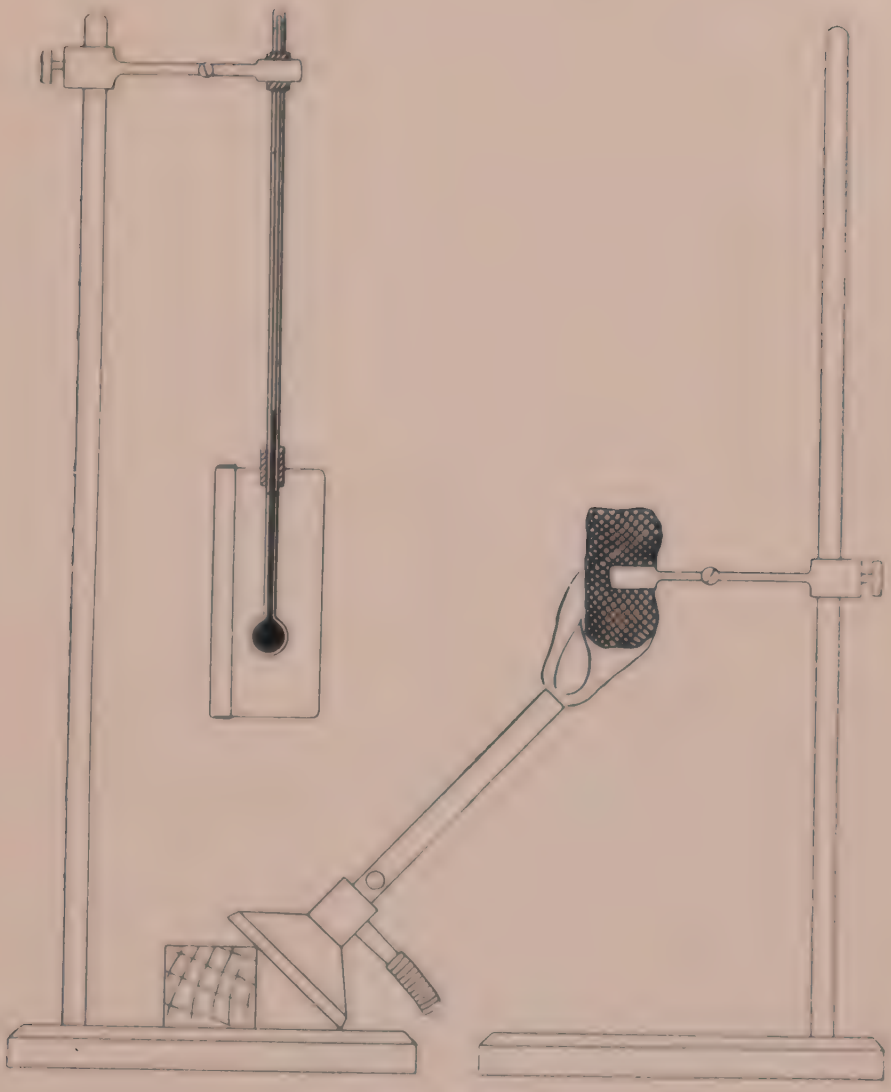


FIG. 26.

Britain white clothes are often worn in summer and darker clothes in winter.

A heat-tight vessel must guard against loss by checking the three ways in which heat leaks, viz.

1. Conduction;
2. Convection;
3. Radiation.

The very best example of a heat-tight vessel that we have is the Vacuum Flask. In the advertisements it is credited with keeping hot liquids hot, i. e. heat inside cannot get out; and cold liquids cold, i. e. heat outside cannot get in.

It was invented by James Dewar and should be known as the Dewar Flask.

The Dewar Flask consists of a *double-walled* glass vessel, the space between the two walls being as nearly as possible a perfect *vacuum*. The glass walls are *silvered*. For scientific purposes the open end is *loosely* plugged with cotton wool. It is very effective, for

(1) Glass is a bad conductor.
(2) A vacuum is a non-conductor. Why?
(3) Convection currents cannot be set up in a vacuum. Why not?
(4) The silvered glass reflects any radiant heat that may fall on it.
(5) Air is a bad conductor.
(6) Entangled air prevents convection.

James Dewar invented the flask to preserve liquid air, the temperature of which was about -195° C. Liquid air is sent by rail in these flasks with very little loss, although liquid Oxygen boils at -182° C. and liquid Nitrogen at -195° C.

CALORIMETRY

Heat is measured by means of a **Calorimeter** (Lat. *calor*, heat). (Fig. 27.)

A calorimeter consists of a metal vessel, usually of copper, although a small cocoa-tin answers the purpose very well.

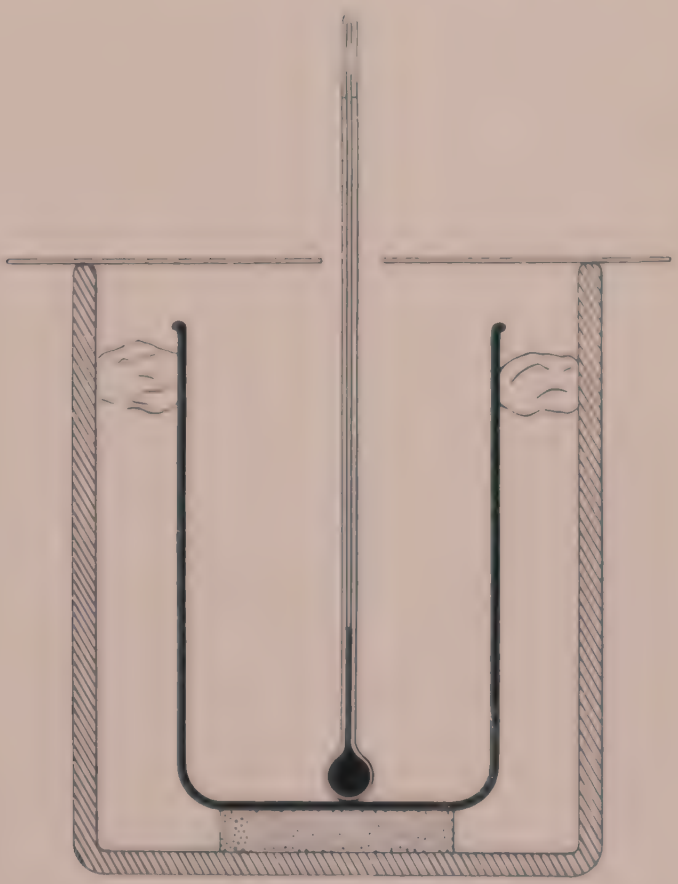


FIG. 27.

(*a*) It should be kept *brightly polished* on the outside. The calorimeter is insulated (Lat. *insula*, an island), as will be seen from the diagram, by (*b*) the *cork* disk on which it rests; (*c*) the cotton wool fitted *loosely* between the rim and the sheath; (*d*) a sheath (an ordinary earthenware jam-jar answers well); and (*e*) a lid of cardboard covering the whole. A thermometer passes through a hole in the cardboard.

Is this apparatus as nearly heat-tight as you can make it? How does it check loss by Conduction, Convection, Radiation? Consider (*a*), (*b*), (*c*), (*d*), (*e*).

The metal vessels should be plainly marked, and a record of the weight of each kept handy for reference. Repeated weighing of the same calorimeters is a waste of time.

THE SIPHON

Before commencing the study of the Physics of the Air, we carried out experiments with water to furnish examples of pressure and pressure changes. Similarly, before commencing the measurement of Heat, we shall consider the action of the siphon, a very useful piece of apparatus in everyday life.

The Siphon. (Gk. *siphōn*, a tube.) Vessels in which water or other liquid is stored are often provided with a tap at the foot—many examples will occur to you. When it is not convenient to fit a tap, a siphon is sometimes used which consists of a tube bent into the form of the letter U or J and inverted (∩ ⅂).

Experiment 50. Take two pieces of glass tubing, bent at right angles, and join them by a small rubber connecting tube as shown in Fig. 28.

Fit a clip to the rubber tube. Fill a gas jar with water, and stand a similar jar containing a little water alongside. Fill the prepared tube with water; clip; close the ends with your fingers; invert; and place one limb in each jar. Open the clip. In which direction does the water flow? When does the flow cease? What causes the water in *X* to flow upwards? Think over the conditions: (1) Air is pressing on the water in each jar, but, by it, the water in *X* is being pushed towards *Y* and the water in *Y* towards *X*. Since these pressures are equal and opposite they neutralize one another.

(2) There is a 'head of water' in each limb of the siphon. The flow ceases when the 'head of water' is the same for each limb. How is 'head of water' measured?

Set the siphon flowing again and clip before the flow ceases. What are the conditions now?

Why does the water not run down from each tube into its own jar? If this happened, what would be left in the top of the apparatus? What is supporting the water columns?

Suppose the tubes were 20 ft., 30 ft., 40 ft. high, would the behaviour still be the same? Is there a height beyond which a siphon will not work? (Bk. I, pp. 85, 86.)

Imagine your apparatus placed under a bell jar from which

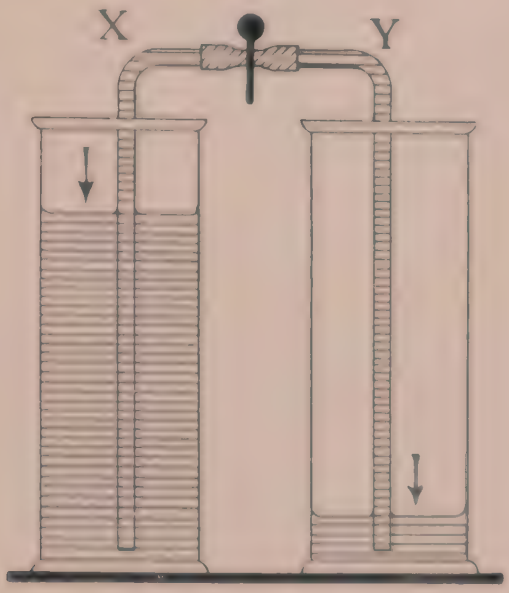


Fig. 28.

all the air had been pumped out, would it still act as a siphon?

The air pressure, then, prevents the water from separating in the tube and leaving a vacuum at the top. The pressure of the water in tube *X*, at the level of the water in its jar, is less than the pressure of the water in tube *Y*, at the level of the water in its jar; because the 'head of water' is less in *X* than in *Y*.

Two equal and opposite pressures produce equilibrium.

Two unequal and opposite pressures produce motion, and the direction of the motion will be the direction of the greater

force. Examples of this are a tug-of-war; unequal weights on the ends of a string over a pulley wheel.

Look at your apparatus, picture to yourself the conditions, open the clip. Write a note in your own words on the action of the siphon.

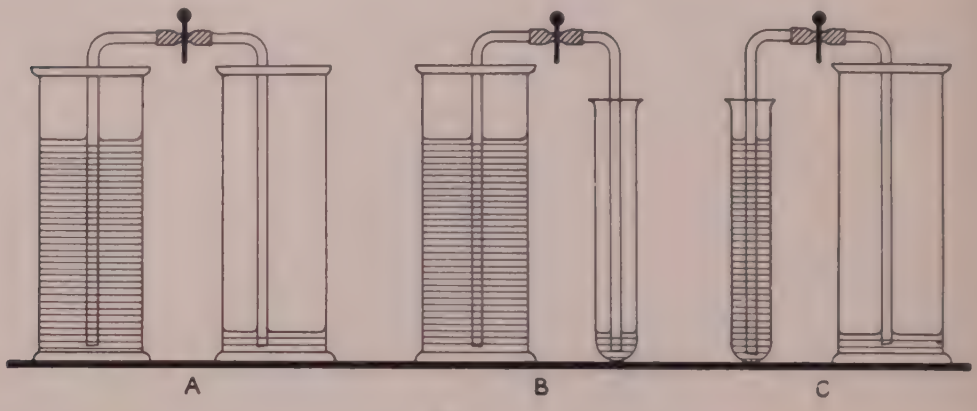


Fig. 29.

Experiment 51. (Three pupils or groups should work together.) Fit up three siphons, using (A) two similar gas jars, (B) a gas jar and a test-tube, (C) a test-tube and a gas jar. (Fig. 29.)

Fill the *left*-hand vessels and put a *little* water in the *right*-hand vessels. Mark the levels with gummed paper strips. Open the clips.

Note:

(1) In each case the flow ceases when the 'head of water' in both limbs is the same.
(2) In case (A) the fall is equal to the rise. Why?
(3) In case (B) the fall is less than the rise. Why?
(4) In case (C) the fall is greater than the rise. Why?

The common level at the end of the experiment, when the vessels are not similar, is always nearer to the initial level of the vessel of larger diameter.

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MEASUREMENT OF HEAT

In dealing with Temperature and the Thermometer we spoke of **Heat-Level**. (Bk. I, p. 39.) Heat flows from a hotter to a cooler body in contact with it until the heat-levels are the same, or, as we say, heat flows from a body at a higher temperature to a body at a lower temperature until the temperature is the same for both bodies. Hot water is cooled by pouring it into a cool vessel, and the vessel is warmed. Why does the housewife warm the teapot?

Experiment 52. Run into the calorimeter from a burette 50 c.cm. water. As a rule in heat experiments it is quite unnecessary to weigh the water used, e. g. 50 c.cm. water at room temperature (say 14° C.) would weigh 49·97 gm. and the burette is a very reliable measuring vessel. The 50 c.cm. may be called 50 gm. Place the calorimeter in its sheath with cork, cotton wool, card, and thermometer.

Take a second 50 c.cm. water in a beaker and heat to 40° C. It is well to heat the water to 45° C. and then allow it to cool to 40° C.

Read the temperature of the water in the calorimeter, pour in the water at 40° C., and stir with the thermometer.

Read the temperature of the mixture.

Record:

50 gm. water at ° C.
50 gm. water at 40° C.

Mixture → 100 gm. water at ° C.

Repeat the experiment, using 20 gm. water in the calorimeter and 80 gm. water at 40° C.

Record:

20 gm. water at ° C.
80 gm. water at 40° C.

Mixture → 100 gm. water at ° C.

Repeat again, using 80 gm. water in the calorimeter and 20 gm. water at 40° C.

Record:

80 gm. water at ° C.
20 gm. water at 40° C.

Mixture → 100 gm. water at ° C.

Note: (1) In each case the rise and fall in temperature cease when a common temperature is reached.

(2) The Common Temperature is about midway between the temperatures of the hot and cold water when the quantity of water is the same.

(3) When different quantities of water are used the Common Temperature is near the initial temperature of the greater quantity.

Compare (1), (2), and (3) with the results of the siphon experiments.

These results fit in with our daily experience.

A pint of boiling water (100° C.) added to a cold bath would have little effect. A gallon of boiling water (100° C.) added to a cold bath would make a great difference. Evidently we cannot say that water at 40° C. mixed with water at 10° C. gives water at $\frac{40° + 10°}{2}$, i. e. 25° C.

That statement would only be true on one condition. What is it?

A large hot-water bottle is much more effective than a small one at the same temperature.

The water required to fill the large bottle would take longer to boil. It would take in more heat from the fire, and therefore have more heat to give out.

This idea of quantity of heat, as distinct from temperature, brings us to a new unit called the **Calorie**.

THE CALORIE AND THE THERM

The Calorie is the quantity of heat required to raise the temperature of one gramme of water through one centigrade degree; or, the calorie is the quantity of heat given out when the temperature of one gramme of water falls through one centigrade degree.

What is the difference between 1° C. and 1 C.°?

Experiment 53. Place a pipe-clay triangle on the tripod stand. Put 100 c.cm. water (i. e. 100 gm.) into a calorimeter, i. e. the metal vessel, and stand it on the pipe-clay triangle. Heat the water, using a Bunsen burner turned down so that the flame just covers the bottom of the calorimeter. Stir gently with a thermometer. By watching the thermometer you will know when the burner has supplied any given number of calories, e. g. 1,000 calories, for:

1 Calorie raises the temperature of 1 gm. water through 1 C.° (Definition).

10 Calories raise the temp. of 1 gm. water through 10 C.°

∴ 10 Calories raise the temp. of 10 gm. water through 1 C.°

1,000 Calories raise the temp. of 100 gm. water through 10 C.°

When the temperature of the 100 gm. of water has been raised through 10 C.° the burner has supplied 1,000 Calories.

Work the following examples:

(1) Find the number of Calories required to raise 3,000 gm. water from 10° C. to 100° C.

1 gm. water is raised through 1 C.° by 1 Cal.

∴ 3,000 gm. water are raised through 1 C.° by 3,000 Cal.

∴ 3,000 gm. water are raised through $(100° - 10°) = 90$ C.° by $90 \times 3{,}000$ Cal.

Answer. 270,000 Calories.

(2) A kettle holds 2,500 c.cm. water. How many Calories would be needed to raise this water from 15° C. to boiling-point?

(3) The water in a hot-water bottle weighs 1,200 gm. How many Calories will it give out in cooling from 90° C. to 40° C.?

1 gm. water falling through 1 C.° gives out 1 Cal., &c.

(4) A Bunsen burner gives 1,200 Cal. per minute. A vessel containing 400 gm. water at 5° C. is put over the burner for $3\frac{1}{2}$ minutes. What will its temperature be?

For the experiments that follow, the number of calories given per minute by the Bunsen burner must first be found.

Experiment 54. To find the number of Calories given per minute by a Bunsen burner.

One pupil, who must be provided with a watch, gives a signal, by tapping the desk, every time the second hand points to 60 and 30, i. e. every half-minute.

It is very helpful if this pupil calls 'Ready' about three seconds before the signal, so that those carrying out the experiment may be ready to read the thermometer.

Run into the Calorimeter from a burette 100 c.cm. water. Stand the Calorimeter on a pipe-clay triangle which rests across a tripod stand. Lower the flame so that it does not spread beyond the rim of the base of the Calorimeter. Stir the water very gently with a thermometer. Take the first reading on the thermometer at a signal when the temperature is 20° C. to 25° C. Take readings every half-minute until the temperature reaches 60° C. to 65° C. Stir gently before each reading.

Record.

	Time.	*Temperature.*
1st reading	0	°C
2nd ,,	½ min. . . .	°C.
3rd ,,	1 min. . . .	°C.
4th ,,	1½ min. . . .	°C.
5th ,,	2 min. . . .	°C.

Total rise in temperature . C.°
Time taken . . . . 2 min.
Average rise per min. . . C.°
Weight of water . . . 100 gm.

Calculate the number of Calories supplied per minute.

Coal-gas on entering a house is measured by the gas-meter in cubic feet. Until recently the gas engineers were only concerned with the number of cubic feet supplied, and the householder was charged accordingly, e.g. 3*s.* 4*d.* per 1,000 cub. ft. The gas-meter, however, gives no information as to the quality of the gas and the consumer is more interested in the amount of heat or number of calories he gets for his money than in the number of cubic feet of gas.

Gas is still measured in cubic feet at the meter, but the gas engineers must guarantee that 1,000 cub. ft. of gas represents not less than a certain heating value. On burning it must supply a certain number of calories of heat.

Now the calorie is not only a very small unit but it is based upon the gramme and the centigrade degree. In Britain the unit employed by engineers is based upon the pound and the

Fahrenheit degree. It is called the British Thermal Unit (B.Th.U.). The British Thermal Unit is the amount of heat required to raise one pound of water through one Fahrenheit degree. How many calories are there to 1 B.Th.U.? (1 F.° = $\frac{5}{9}$ C.° 1 lb. = 450 gm.)

Gas engineers use a very large unit called a **Therm**, which is 100,000 British Thermal Units. A therm would raise 100,000 lb. of water through 1 F.°, or since 1 gall. of water weighs 10 lb. it would raise 10,000 gallons of water through 1 F.°

In Edinburgh, for example, one cubic foot is guaranteed to give 500 B.Th.U., i. e. on burning to raise 500 lb. or 50 gall. water through 1 F.°, or 1,000 cubic feet is guaranteed to raise 500,000 lb. water 1 F.°, i. e. to provide 5 **Therms**.

If you use gas, look at the meter card and find how much you pay per therm.

LATENT HEAT

The numbers on the scale of the thermometer, as we have seen, depend upon two Fixed Points, viz. the Boiling-Point and the Freezing-Point of Water.

Celsius of Upsala in 1741 suggested that these two points should be taken as 100 and 0 respectively, and so gave us the Centigrade thermometer.

When water is being heated the temperature rises until it reaches the Fixed Point 100° C., and further continued heating causes no change in temperature. (Bk. I, p. 38.)

Experiment 55. Pour a few drops of methylated spirit (or better, ether) on the hand, and note the sensation of cold as it evaporates.

Experiment 56 (Demonstration). Pour a few drops of water on the table and stand a beaker containing a little ether on the wet patch. By means of bellows blow air vigorously through a tube dipping into the ether. The abstraction of the heat required to evaporate the ether will reduce the temperature below freezing-point and the beaker will be firmly

attached to the table by the ice into which the water is converted.

Joseph Black, who became one of the most famous of Edinburgh's Professors of Chemistry, suggested experiments on 'Cold produced by Evaporation' when he was a student in Glasgow. In 1762 he described experiments he had himself performed and used the word **Latent** to signify that the heat appeared to go into hiding when water changed to steam, and when ice changed to water.

The Latent Heat of a substance is the number of calories required to change one gramme of it from the solid to the liquid state or from the liquid to the vapour or gaseous state, without change of temperature.

The Latent Heat of Fusion of Ice, or the Latent Heat of Water, is the number of Calories required to change one gramme of ice at 0° C. into water at 0° C.

Experiment 57. Find the number of calories given per minute by your burner. (Make the flame smaller than in the steam experiment.)

Dry the calorimeter, break ice into quite small pieces, and dry them with a towel. Fill the calorimeter with pieces of dry ice. Place it over the flame for three minutes. Pour the resulting water into a burette, using a funnel provided with a piece of wire gauze to keep back any pieces of ice. Read the number of c.cm. of water obtained from the ice melted in three minutes. (If all the ice is melted, repeat the experiment, taking two minutes.)

Record.

The Bunsen supplies Cal. per min.
∴ Calories supplied in 3 min. = Cal.
∴ gm. ice were melted by Cal.
∴ 1 gm. ice at 0° C. required Cal. to change it to water at 0° C.

Very accurate experiments give 80 Cal.

The Latent Heat of steam is 536 Cal., i. e. 536 Calories of Heat are required to change 1 gm. water at 100° C. to 1 gm. steam at 100° C.

The Latent Heat of water is 80 Cal. . . . Complete this, using the preceding as a pattern.

The Latent Heat of all substances that melt or boil can be found. Whenever there is a change of state, heat is either required or liberated: required to change from solid → liquid → gas; liberated to change from gas → liquid → solid.

The Latent Heat of Vaporization of Water or the Latent Heat of Steam, is the number of calories required to change one gramme of water at 100° C. into steam at 100° C.

Alcohol boils at 78° C. Make a similar statement for alcohol.

Experiment 58. Find, as in Experiment 54, the number of calories supplied per minute by your Bunsen burner. See that the flame is adjusted properly, and do not touch the gas tap after you have once started the experiment. Stop the readings before passing 70° C. and remove the thermometer.

Watch the surface of the water for the first sign of boiling and note the time. Allow the water to boil for six minutes.

Cool by standing the calorimeter in cold water. Measure the remaining water by pouring it back into the burette. See that the water stands at '50' before pouring any in.

Record, as in this example.

The Bunsen supplies 1,250 calories per minute.

∴ Calories supplied in 6 minutes = 7,500 Cals.

Water taken	100 gm.
Water left	85 gm.
Water evaporated	15 gm.

∴ 15 gm. water were evaporated by 7,500 Cal.

∴ 1 gm. water at 100° C. required $\frac{7500}{15}$ or 500 Cal. to change it to steam at 100° C.

Very accurate experiments give 536 calories. Note that it requires 536 times as much heat to change 1 gm. water at 100° C. to steam at 100° C. as it does to heat 1 gm. water through 1 C.°

When the steam condenses to water these calories are released, so that 1 gm. steam at 100° C. would lose 536 calories on changing to water at 100° C.

A scald from the steam of boiling water is much worse than a scald from boiling water.

Steam is used at public washhouses to heat the water in the wash-tubs.

Steam is used in restaurants to heat beverages.

Water coolers are used in tropical countries. They are bottles made of porous earthenware so that when filled there is always a film of water on the outer surface. This film evaporates and heat taken from the bottle is used to change the water into vapour, just as, in the ether experiment, heat is taken from the hand. The heat 'hidden' in the vapour is taken from the bottle.

Butter is kept cool in hot weather by placing it in a basin that stands in a soup-plate filled with water. An inverted flower pot is placed over the butter with its rim in the water. The flower pot is thus kept damp and the water evaporating from its surface lowers the temperature.

A dog's nose always seems cold. Why?

HYGROMETRY

Experiment 59. Take two similar thermometers and hang them side by side. Wrap a piece of muslin round the bulb of one and allow the end to dip into a dish of water. (Fig. 30.) Take the readings of the two thermometers every day for a fortnight and record them, noting also whether the day is dry, damp, or wet. Are the roads 'drying up'? Such a pair of thermometers is known as a 'Wet- and Dry-bulb Thermometer'.

It is a wet **Hygrometer**. (Greek *hugros*, wet.)

Date.	*Wet-bulb Thermometer.*	*Dry-bulb Thermometer.*	*Weather note.*
	°C.	°C.	

On some days the readings are almost alike, on others there may be a difference of several degrees. As we saw in the ether experiment, rapid evaporation produces great cold. Evidently, then, on some days evaporation of water takes place more rapidly than on others. Look at your record and state what connexion there is between the readings and the

Fig. 30.

weather. Why is the reading of the wet-bulb thermometer never higher than the reading of the dry-bulb thermometer?

When would you advise a laundrymaid to hang clothes out to dry; on a day when the thermometers are alike or on a day when they differ considerably?

When the thermometers are alike, walls, the pavements, iron railings, &c., are damp. Explain.

FREEZING AND BOILING

Freezing and boiling are so familiar that it will be worth while to consider them more closely.

Whether water is solid, liquid, or gas depends upon temperature. At and above 100° C. it is gaseous.

From 0° C. to 100° C. (both inclusive) it is liquid.

At and below 0° C. it is solid.

100° C. is both Boiling-Point and Condensing-Point.

0° C. is both Melting-Point and Freezing-Point.

Freezing.

Melting and freezing are not confined to water.

Butter is often placed near the fire in winter to soften it and, if forgotten, soon melts to a yellow oil.

The wax of candles melts near the flame. Only by becoming liquid is it possible for it to pass up the wick.

Sulphur melts on heating. (Bk. I, p. 32.)

Many metal articles are cast, i. e. made by melting the metal and pouring it into a mould.

Experiment 60. Take a bottle (a medicine bottle) and *fill* it with water. Cork the bottle and tie the cork firmly. Stand the bottle in a freezing mixture (salt and pounded ice) and leave it until the water is completely frozen. Note what has happened.

An interesting experiment was performed many years ago at Quebec by Major Williams, who filled a twelve-inch iron shell with water, plugged the hole with a wooden stopper driven in with a mallet, and left it outside on a frosty night. The temperature fell to − 18° F. The wooden plug was projected over 100 yards and a cylinder of ice 8 inches in length protruded from the hole.

To our cost, we unwittingly allow a similar experiment to take place at home when an unprotected water-pipe is frozen. The thaw reveals the mischief done.

Water on freezing expands, and the volume of the ice is greater than the volume of the water.

1 gm. water must give 1 gm. ice (The Law of the Conservation of Matter);

but 1 c.cm. water gives 1·1 c.cm. ice,

or 100 c.cm. water gives 110 c.cm. ice,

and ∴ 1 c.cm. ice weighs $\frac{1 \text{ gm.}}{1 \cdot 1} = 0 \cdot 91$ gm.

Explain (1) why pieces of ice float in water;

(2) why ice forms on the *surface* of pools, &c.;

(3) Icebergs.

Experiment 61. Half fill a test-tube with paraffin wax. Melt the wax by standing the test-tube in a water bath. Stand it on one side to freeze, and note the surface of the solidified wax carefully. Does liquid wax change its volume on solidifying? Does it change its density? Will solid wax float or sink when dropped into liquid wax?

Melt the wax again and drop into it a small piece of solid wax. Does it behave as you expected?

Boiling.

Sir Francis Younghusband in his 'Epic of Mount Everest' says: 'At these high altitudes (26,000 ft.) water boils at so low a temperature that a really hot drink cannot be had. Nothing more than tepid tea or soup can be concocted.'

At Quito water boils at 90° C. Where is Quito?

On the top of Mont Blanc water boils at 84° C.

Pascal argued that the air pressure would be less at the top of mountains. (Bk. I, p. 85.)

The following experiment was devised by Benjamin Franklin (1706–1790).

Experiment 62. Half fill a *round-bottomed* flask with water. Boil the water vigorously for a few minutes so that the steam may expel the air in the flask. (See Bk. I, p. 50.) Remove the Bunsen burner and immediately insert a rubber stopper through which passes a thermometer. Invert the flask as shown, and pour cold water over it. Notice what happens, and read the thermometer. (Fig. 31, p. 79.)

What happens to the vapour in the flask when the cold water is poured over it?

This experiment may be repeated several times without re-heating.

Record the lowest temperature at which boiling takes place.

What connexion is there between this experiment and the experience of mountaineers?

How would the boiling-point of water be affected at the bottom of a deep mine?

Water vapour is always present in the air and is *invisible*. When a tumbler of cold water is brought into a warm room beads of moisture form on the glass; in a crowded room the windows, especially on cold nights, are 'steamed'; if we breathe on a mirror it is dimmed. In none of these cases was the water visible until the vapour condensed on a cold surface. Fog is not water vapour: it is condensed vapour, i.e. tiny drops of water. When water or other liquids evaporate in a closed vessel the invisible vapour behaves like a gas, and exerts pressure on the walls of the vessel, and on the surface of the liquid.

Experiment 63 (Demonstration). Fill a Torricellian tube with mercury. Place the finger over the open end and invert the tube over a dish of mercury. Remove the finger when the end of the tube is below the surface of the mercury, to ensure that no air enters the tube. Record the 'height of the barometer'. (Fig. 32.)

By means of a modified fountain-pen filler (Fig. 32*b*) drive a small quantity of ether into the tube.

Record the 'height of the barometer'. Note the change. What pressure is the ether vapour exerting on the mercury in the barometer? Give the result in mm. of mercury. Wring a cloth out of boiling water and wrap it round the upper part of the barometer.

Record the 'height of the barometer', and calculate the pressure of the ether vapour.

What is the effect of temperature on vapour pressure?

Steam-engines make use of vapour pressure. The gauge may register 100 lb., which means that the vapour pressure in the boiler is 100 lb. per sq. in. The water in the boiler would be boiling at a temperature of 150° C.

Water, or other liquid, boils when its vapour pressure is

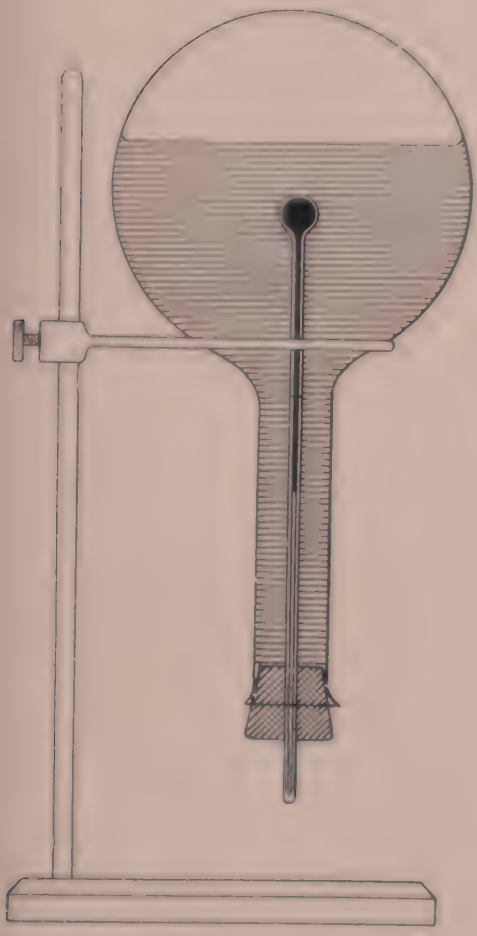

FIG. 31.

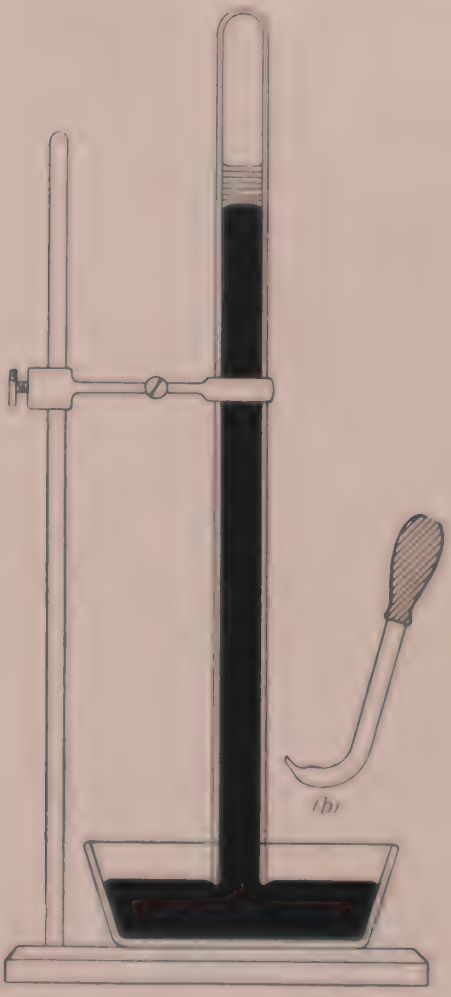

FIG. 32.

great enough to overcome the pressure on its surface, for only then can the vapour escape freely.

Water boils at 100° C. when the barometer stands at 760 mm. or 30 in., i. e. when the pressure is 14·7 lb. per sq. in.

Water would boil at 0° C. if the pressure were reduced to 4·6 mm. or 0·18 in. of mercury, i. e. when the pressure is $1\frac{1}{2}$ oz. per sq. in.

In Britain the barometric readings vary between 28 in. and 31·5 in. The following table shows how the boiling-point of water is affected :

Pressure of the Atmosphere.	*Temperature of Boiling Water.*
28 in.	98° C.
29 in.	99·1° C.
30 in.	100° C.
31 in.	100·9° C.
31·5 in.	101·5° C.

In experiments the thermometer in boiling water sometimes stands steadily at 99° C. or 99·5° C. The table given explains this.

SPECIFIC HEAT

In the definition of the Calorie, water is specially mentioned ; just as it is in the definition of the Gramme, the fixing of the scale on the Thermometer, and the definition of Specific Gravity. Water is taken as the standard substance, and if any other substances were substituted for it, a new unit would appear, and all the results founded on the old one would be changed, e. g. if the weight of 1 c.cm. of Gold were taken as the unit, the new gramme weight would be as large as the present 20 gm. weight.

The question then arises, 'Does it take exactly as much heat to raise the temperature of 1 gm. of other substances through 1 C.° as it does to raise the temperature of 1 gm. of water through 1 C.° ?'

Experiment 64. Find the number of calories given per minute by the Bunsen burner.

Take 100 gm. lead shot (weigh to the nearest shot and do not use the small weights). Put 100 gm. (c.cm.) water in the calorimeter. Add the shot, and find the average rise in temperature per minute when the mixture is heated over the Bunsen burner.

Record.

No. of Cal. given per min. = Cal.
Rise in temperature of mixture of water and lead per min. = C.°
1 gm. water raised 1 C.° uses 1 calorie.
∴ The 100 gm. water raised C.° used Cal.
∴ The 100 gm. lead raised C.° used the remaining Cal., viz. Cal.

Was the heat supplied divided equally between the water and the lead?

How many calories would be required to raise 1 gm. of lead 1 C.°?

Experiment 65. Repeat the experiment, using Aluminium (pieces of thick wire or boot tackets), Iron (cycle balls), Copper (rivets), Glass (beads), and collect class results.

One point stands out very clearly, viz. that taking the quantity of heat required to raise 1 gm. water 1 C.° as 1 Calorie, for other substances it is much less than 1 Calorie.

Water is said to have a great heat capacity (Lat. *capax*, roomy), a great roominess for heat. It can hold a great deal; compared with other substances it requires much heat to raise its temperature.

In Specific Gravity water is taken as the standard; its s.g. is 1. The s.g. of turpentine is 0·8, which means that when equal volumes of water and turpentine are taken the turpentine will weigh 0·8 of the weight of the water.

Specific Heat

When we consider the amount of heat or number of calories required to raise the temperature of a given weight of a substance through a certain range of temperature we always compare it with the amount required to raise the same weight of water through the same range, e. g.

Amount of heat required to raise 1 gm. Water 1 C.° =	1 Cal.
Amount of heat required to raise 1 gm. Lead 1 C.° =	0·03 Cal.
Amount of heat required to raise 1 gm. Mercury 1 C.° =	0·03 Cal.
Amount of heat required to raise 1 gm. Copper 1 C.° =	0·09 Cal.

The Specific Heat of a substance is the number that expresses what fraction of the heat used by water in being raised through a given range of temperature is required to raise the same mass of the substance through the same range.

$$\text{S.H.} = \frac{\text{Cal. required to raise any mass of the substance through any range of temperature.}}{\text{Cal. required to raise the same mass of water through the same range of temperature.}}$$

$$\text{or} = \frac{\text{Cal. required to raise 1 gm. substance 1 C.°}}{\text{Cal. required to raise 1 gm. water 1 C.°, viz. 1 Cal.}}$$

or for all practical purposes

$$\text{S.H.} = \frac{\text{Cal. required to raise 1 gm. substance 1 C.°}}{1}$$

thus: S.H. of copper is approximately $\frac{1}{10}$, means that any mass of copper would require only $\frac{1}{10}$ of the heat to raise it through any range of temperature that the same mass of water would require for the same range. Imagine 1 lb. of water and 1 lb. of copper both at 10° C. placed in an oven together. Then, when the water reached 12° C., the copper would be at 30° C. The heat that raises water 2 C.° raises the same mass of copper 20 C.° because copper only requires $\frac{1}{10}$ of the heat for the same effect.

A Table of Specific Heats.

Water	1
Paraffin	0·5
Turpentine	0·46
Aluminium	0·2
Glass	0·16
Iron	0·11
Copper	0·095
Mercury	0·033
Lead	0·032

Water is used for storing heat because of its great capacity for heat. A large quantity of heat is required to make it hot, and therefore it has a large quantity to give out during cooling. The water in a hot-water bottle would provide much more heat than the same weight of any other substance at the same temperature.

The sea warms slowly in the summer because it requires so much heat to raise its temperature. It has a good deal of heat to lose in the winter, and therefore cools slowly. Countries surrounded by sea have a Maritime Climate, i. e. not very hot in summer and not very cold in winter. Countries far away from the sea, like Central Russia, have an Extreme Climate. The land is warmed more quickly than water in the summer, because the specific heat of soil and rock is only about 0·2, but for the same reason it cools more quickly in the winter.

Hot-water heating systems depend on Convection Currents, as we have seen, but they are very effective because the heated water conveys so much heat from the furnace owing to its great heat capacity.

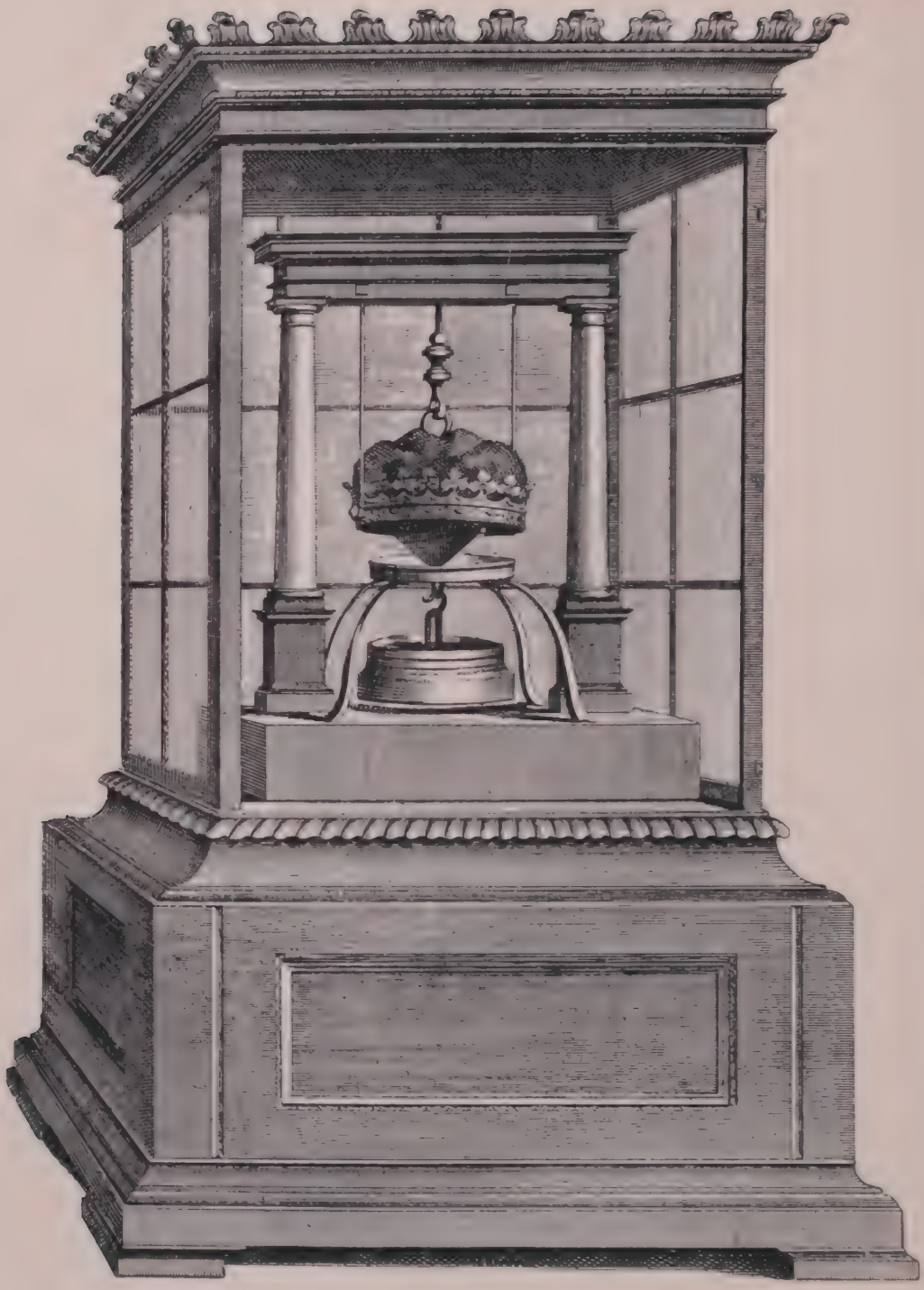

Fig. 33.

III

MAGNETISM

PROPERTIES OF MAGNETS

MOST children have played with a horseshoe magnet, and know the game of fish-ponds in which metal fish are caught with an iron hook in an apparently magical manner.

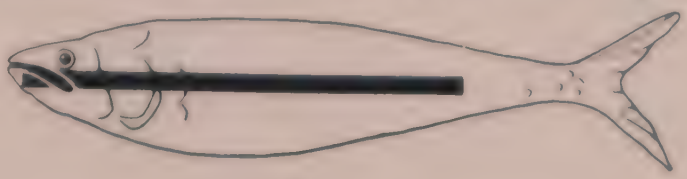

FIG. 34.

If the fish were made of glass they would appear as in Fig. 34, and reveal a small magnet within. Very many useful devices and machines met with in everyday life depend upon this same principle; viz. that magnets attract iron.

Something was known of magnetism 4,000 years ago by the Greeks and the Chinese. Reports of events that happened in the far distant past are not very reliable, and the story that a shepherd in Asia Minor, named Magnes, found it difficult to remove his crook, shod with iron, from a dark rock against which he had placed it, must probably be taken *cum grano salis*. One thing is certain: pieces of a dark rock found in Magnesia, Asia Minor, do attract small objects made of iron. The rock is an iron ore composed of Iron and Oxygen, and is called Magnetite, Magnetic oxide of Iron, or Natural Magnet.

Experiment 66. Have ready on a sheet of paper a small heap of coarse iron filings and as the magnetite is passed to you, examine it. Cover it with iron filings and then lift it carefully. Sketch the magnetite with the filings adhering to it, showing clearly where they appear to be most crowded.

There is a famous natural magnet at the Oxford University Museum which was presented by the Countess of Westmorland in 1756. The weight shown in the figure is 160 lb., but at a recent test it was found that 170 lb could be held. (Fig. 33.)

Another property of magnetite was also discovered by the ancients.

Experiment 67 (Demonstration). Suspend the piece of magnetite by means of unspun silk from a wooden stand. Why not use an iron stand? Tie the silk midway between the points at which the filings were most crowded in Experiment 66.

When the magnetite comes to rest, mark on the table the direction in which it is pointing. Disturb it and again note its direction when it comes to rest.

A natural magnet, allowed to swing quite freely, always sets in a north-south direction. This property was used by the Chinese, who are said to have taken magnets to guide them across the Plains of Tartary and the Indian Ocean, 3,000 years ago. The name Lodestone, meaning Leading Stone, is often given to Magnetite. (Cf. Lodestar. Which is the Lodestar ?)

At an early date it was discovered that a natural magnet, placed in contact with iron, handed on its property of attracting small pieces of iron without losing any of its own magnetism. It was thus possible to produce *Artificial Magnets*. These are always used for experiments, for they can be made in convenient shapes and sizes; and are also much stronger than natural ones.

Steel and Soft Iron.—Consider a knitting-needle, sewing-needle, pen-point, knife-blade, clock-spring. Notice that all these are made of tempered steel which is hard and brittle. If you attempt to bend any of them they snap suddenly. Broken needles, pen-points, knife-blades, are very common. Contrast these with a strip cut from a cigarette-tin or cocoa-tin. (Called *tin*, because the iron from which it is made has been coated with tin. Can you suggest why the iron is coated in this way?)

When iron, that bends easily and is not brittle, is intended in magnetism, it is termed *Soft Iron* to distinguish it from tempered steel, usually termed *Steel*.

Experiment 68. Take any piece of steel—a knife-blade or pen-point will do—and place it on the table. Stroke it ten times with *one end* of a bar magnet, and in *one direction*, lifting the bar magnet well away on the return journey. (Fig. 35.) The dotted lines in the figure indicate the path of the magnet.

This is known as magnetizing by *single touch*. Test the piece of steel for magnetism by bringing it near small articles made of iron such as gramophone needles, iron tacks, iron filings.

Experiment 69. Take two pieces of clock-spring (4 in. long), heat them to redness in the Bunsen flame. When they are bright red, close the air holes of the burner and turn the gas

Fig. 35.

down gradually, so that the strips cool slowly. Hammer the strips flat and make a notch at one end of each with a file. You will find that the iron is not nearly so brittle as before. Harden it again by heating to bright redness and plunging suddenly—while still red-hot—into cold water. You may test an odd piece and note how it has regained its brittleness. Magnetize one strip by the method of single touch, using the marked end of the bar magnet and commencing at the notch on the strip. (The mark may be a line scored across or the letter N.)

Show exactly how you proceeded by a sketch, similar to Fig. 35.

Test your strip by the methods of Experiments 67 and 68, and record the results.

It is somewhat tiresome to tie a strip so that it balances well. Here are three ways of overcoming the difficulty. (*a*) Make a small paper or wire stirrup. A piece of paper 2 in. × ½ in. is doubled lengthwise, and a small hole is made to take the silk ; or (*b*) a piece of brass or copper wire 4 in. long is first doubled like a hair-pin, and then the ends are turned up ; or, better, (*c*) push the eye-end of a darning-needle into a flat cork and place over the needle an inverted test-tube 2 in. × ½ in. Fix your magnetized strip to the test-tube with a

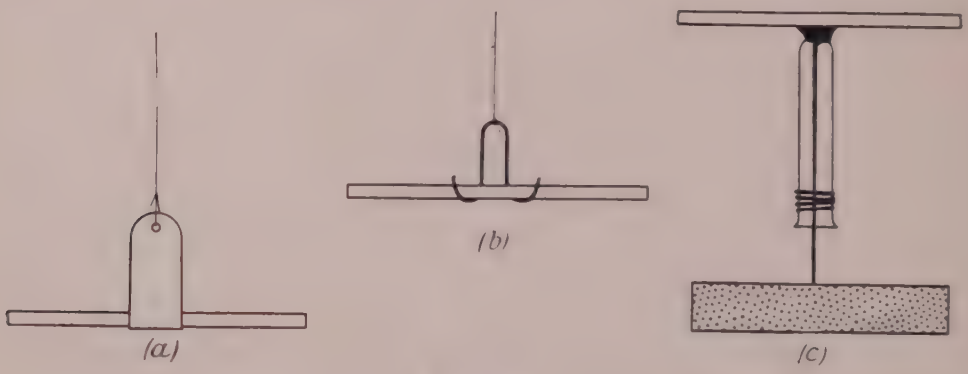

Fig. 36.

small piece of plasticene. If this requires steadying, wind a wire collar round the test-tube. (Fig. 36 (*a*), (*b*), (*c*).)

Note that the artificial magnet also attracts iron filings and that they adhere chiefly at the ends. The ends of a magnet where the magnetism appears to be strongest are called **Poles.**

The Pole which points N-ward is the **North Pole** of the magnet, usually marked **N.** The other Pole is the **South Pole** of the magnet and is not marked.

A magnet supported so that it is quite free to turn is known as a magnetic needle. (Fig. 36.)

Experiment 70. Take the second strip and magnetize it by single touch, this time using the unmarked end of the bar magnet and commencing at the unnotched end of the strip. Make a sketch to show the method.

Support, as before, so that it may turn freely.

Note again that the notched end is the N-pole of the magnet. Be quite sure before beginning the next experiment that the notched ends of your two magnets are the N-poles. This is very important.

Experiment 71. (*a*) Support one of the magnets, as before, and, when it has come to rest, bring the N-pole of the other magnet slowly towards its N-pole and watch its behaviour. (*b*) Take your magnet away, and when the magnet comes to rest bring the S-pole up to the S-pole.

Record the results.

(*a*) A N-pole repels a N-pole.

Write the corresponding statement for (*b*).

This may be summed up thus:

Like Poles Repel.

Experiment 72. Find the effect when *unlike* poles are brought together. Record and sum up as in Experiment 71.

Experiment 73. Cut a similar strip from a piece of cocoa-tin, and set it swinging freely. Try the effect of the N-pole of a magnet on each end. Repeat with a S-pole. Record. How does the behaviour of this unmagnetized strip differ from that of the magnetized strip?

THE FIRST LAW OF MAGNETISM { *Like Poles Repel.* / *Unlike Poles Attract.*

What is a Law?

Note that magnets *attract* iron, i. e. unmagnetized iron; and unlike poles *attract* one another.

What, then, is the test for determining whether a strip of iron has been magnetized?

Scientists always try to give a reasonable explanation for phenomena. They make a guess called a theory, and state that, at any rate, if the guess is correct the results observed would follow. Unfortunately, when objections are raised the originator and his followers sometimes try to make the *facts* fit the *theory*.

Here is an example from Chemistry.

As we have already seen, the early chemists were very

anxious to explain combustion, e.g. Jean Rey, Lavoisier, Priestley.

Stahl (1660–1734) imagined that combustible substances contained something which escaped when they burned. To this, in 1723, he gave the name Phlogiston.

A Fact: Substances could not burn in an enclosed space.

Stahl's Theory: This was explained by saying that the Phlogiston could not escape. (How would you explain it?)

A Fact: Substances became heavier when they were burned. (Give examples.)

Stahl's Theory: This was awkward, but an explanation was forthcoming. Phlogiston must be so very light that it weighs less than nothing, and therefore a substance, before the Phlogiston had escaped, must be lighter than it is after the Phlogiston is away.

(How do you explain the increase in weight on burning?)

This theory was accepted as the true explanation of combustion for half a century. Lavoisier's experiments, 1774, dethroned it. (Recall them.)

We smile at Phlogiston to-day but, after all, the only way to make progress is to make trials and then seek to discover whether they are in the right direction or not. This theory set a number of experimenters at work, some of whom were trying to demonstrate its truth and others who were just as eager to show that the explanation was not sound.

Scientists very naturally endeavoured to imagine what was going on in the piece of steel as it was being magnetized. Some pictured the metal as made up of an immense number of very small particles. You have handled iron filings, small particles that originally made up a bar of iron. Each of these filings could again be reduced to smaller filings and in imagination we can go on thinking of them smaller and smaller still. To these smallest imaginable particles of iron the name Molecule was given (Lat. *moles*, mass; *cule*, a diminutive ending. Compare English, lamb-*kin*) and the theory was called the Molecular Theory of Magnetism.

A steel bar, then, is supposed to be made up of innumerable particles or molecules, each one of which is a tiny magnet. As these are all jumbled together they neutralize one another. On stroking (Single Touch) the bar with the S-pole of an artificial magnet the unlike or N-poles of these tiny magnets are attracted, many of them turn and face in the same

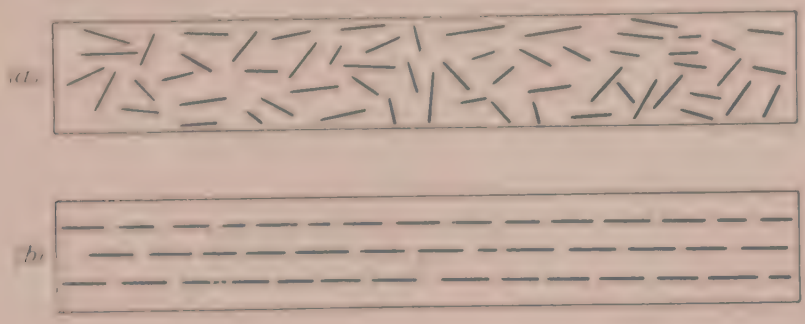

Fig. 37.

direction and the steel bar shows this by acting as a magnet. (Fig. 37 (*a*), (*b*).)

Experiment 74. Take a magnetized strip and break it into two or more pieces. Test each of the pieces. You will find that they are all complete magnets possessing two poles. In fact it is impossible to make a magnet with one pole only. Look at Fig. 37(*b*). The Molecular Theory explains this satisfactorily.

Consider the following:

John Prescott Joule (1818–89) showed that a bar of steel increases slightly in length when it is magnetized.

If a magnet is roughly handled it becomes weaker.

It has been found that if a bar of steel is slightly magnetized and then put into nitric acid until the outside layer has been dissolved away it loses its magnetism. Does the theory account for these facts?

Experiment 75. Take a test-tube and fill it with iron filings. Hold the end near a suspended magnet, first to the N-pole and then to the S-pole. What do you conclude? Lay it on the table and proceed as you did in Experiment 68, i. *e.* magnetize it by single touch. Watch the movement of the filings. Lift the test-tube very carefully and again hold the end near the

poles of the suspended magnet. Do you agree that there is now evidence of magnetism? What is the evidence? Does the Molecular Theory appear to be a reasonable explanation?

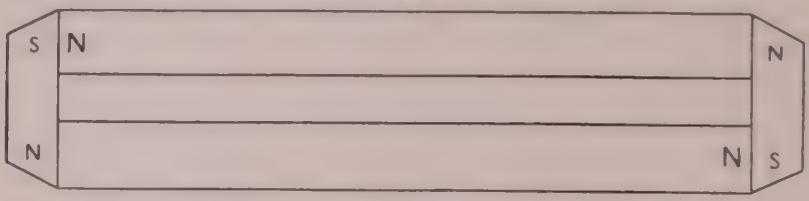

Fig. 38.

What would you expect to happen if you were to magnetize a strip by beginning at the middle and stroking first to one end ten times, and then to the other with the same pole of the magnet? Sketch, and then try it.

Magnets are always supplied with small bars of soft iron in contact with their poles, as shown in the figure. (Fig. 38.)

These are called *keepers*. Magnets should always be stored with the keepers in position.

What happens when soft iron is placed in contact with a magnet? The next experiment will help you to answer the question.

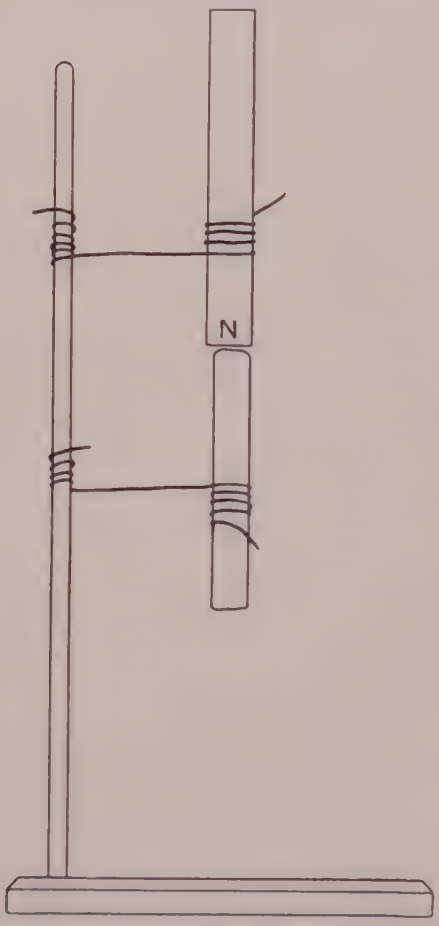

Fig. 39.

Experiment 76. (*a*) Take a piece of soft iron (cocoa-tin) about 4 in. or 5 in. long, or a large nail, and fasten it vertically to a wooden stand with copper wire. Fix a fairly strong bar magnet above with its N-pole touching the strip. (Fig. 39.)

Bring up to the lower end of the strip some iron filings. When you are satisfied that the strip acts as a magnet, remove the bar magnet carefully and note what happens to the iron filings.

Replace the soft iron by a piece of steel clock-spring, and repeat the experiment.

What difference do you suppose there is between the condition of the molecules in soft iron and in hard steel?

(*b*) Repeat the experiment and find, by testing, how the poles are arranged. Now reverse the bar magnet so that its S-pole touches the strip, and test the poles. Make sketches showing the results.

Use the Molecular Theory to explain them.

What happens to the soft iron keepers when they are in position?

How do they help the magnets to *keep* their magnetism? (Fig. 38.)

Magnetism that appears in iron which is near a magnet is said to be **Induced** magnetism. The iron is magnetized by **Induction.** (Lat. *duco*, lead; *in*, in.)

When soft iron is magnetized by induction the magnetism is only present so long as the inducing magnet is there. Whenever the magnet is removed the soft iron loses its magnetism.

Soft iron becomes a **Temporary Magnet** by Induction.

Hard steel becomes a **Permanent Magnet** by Induction.

Every artificial magnet is an example of Permanent Magnetism.

The space in the neighbourhood of a magnet, throughout which the influence of the magnet is felt, is called the Field of the Magnet. The force in the Field due to the magnet acts along well-marked lines. These **Lines of Force** can be shown by experiment.

Experiment 77. Place a bar magnet flat on the desk and lay a sheet of paper over it. Sprinkle iron filings thinly over the paper and tap the paper gently with the end of a pencil. The filings will arrange themselves along lines of force in the field of the magnet. Make a sketch showing the lines of force.

Experiment 78. Push one end of a magnetized sewing-needle into a small piece of cork and place it in a trough of water. It must float vertically with the N-pole *just* above the water.

Support a bar magnet so that it is lying on the surface of the water. (Fig. 40.) Bring the floating needle near the side of the N-pole of the magnet, and note how the needle travels along a **Line of Force.**

Terrestrial Magnetism (Lat. *terra*, the earth).—We have already spoken of the way in which the Chinese used magnets to guide them over trackless deserts and seas, and we have had many examples, in the experiments, of magnets setting North and South when allowed to turn freely.

The world itself, therefore, acts on suspended magnets like a huge magnet with its poles *near* the North and South Geographical Poles. Columbus discovered that the Magnetic Pole was not at the Geographical Pole. (Which one, N. or S. ?) How did he find this out without going there ? Could you draw two lines, one pointing to the Geographical and the other to the Magnetic N-pole ?

In 1831 James Ross discovered the North Magnetic Pole in the north of British North America. Lat. 73° N., Long. 96° W. In 1909 Shackleton found the South Magnetic Pole in the Antarctic Circle. Lat. 72° S., Long. 155° E.

Look these positions out on your map.

Can you suggest a method of finding out when you are at the N-magnetic pole ? What would a magnetic needle do if you carried it away from the Pole ? Would the direction in which you walked make any difference ?

Since the end of a magnet that points to the North has been given the name N-pole the polarity of the earth's magnetism is a little confusing.

The Earth's Magnetic Pole, near the N-Geographical Pole, attracts the N-pole of the suspended magnet and by the First Law of Magnetism must therefore be acting as a S-pole. What is the First Law ?

Some have suggested the name N-seeking Pole for the end of a magnet which seeks the North, but it has been found more convenient to call it the N-pole.

It is interesting to see one of the effects of the earth's

magnetism. If you take a small pocket-compass such as is used for finding direction—often they are made to hang on a watch chain—and hold it close to the foot of an iron bed post you will find that the bed-post is a magnet with a N-pole at the end on the ground. In New Zealand the foot of an iron bed-post is always a S-pole.

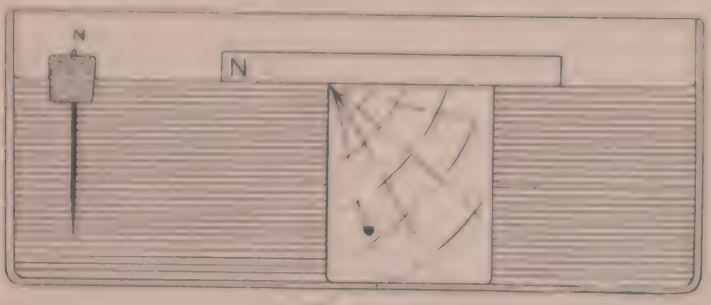

Fig. 40.

Fig. 41.

Experiment 79. Take an ordinary iron poker and hold it in the hand, almost upright (about 70°), with the lower end pointing N-wards. Tap the top of the poker with a hammer a few times and then test the lower end of the poker for magnetism.

The foregoing are examples of magnetization by the Earth's Induction.

THE MARINER'S COMPASS

The Mariner's Compass consists of a card, thin metal or mica disk, to the back of which a ladder of eight magnetized needles is fixed. The card is pivoted at its centre on an accurately made agate bearing. The top of the card shows the *points of the compass.* The whole rests in a basin filled with liquid which prevents the card from rocking and also brings it to rest quickly. The basin is weighted and supported on gimbals so that movements of the ship are not communicated to the compass card. (Fig. 41.) The Mariner's Compass was greatly improved by Lord Kelvin (1824–1907), who did much to make it a sensitive instrument. He introduced the ladder of eight needles. When ships began to be made of iron, the effect of this metal on the compass needles had to be considered. The Earth's Induction caused much of the iron used in the construction of the ship to be magnetic. The influence of this is corrected by placing bar magnets below the compass and soft iron spheres alongside.

STANDARD MARINER'S COMPASS.

THE ELECTROMAGNET

Experiment 80. Take a beaker partly filled with very dilute Sulphuric Acid. Put a strip of Copper (5 in. × 1½ in.) into the acid and observe it carefully. [It is advisable to stand the beaker in a pneumatic trough as Sulphuric Acid damages the clothing.] Remove the Copper, and put in a Zinc strip. You should not be surprised at the result (p. 12). Explain.

Now put in both strips at the same time, keeping them well apart. Watch the strips as you bring the lower ends in contact with one another.

Wind a piece of *bare* Copper wire round the top of each strip. Put the strips in the acid—keeping them apart—and hold the free ends of the Copper wires together. Note what happens in the beaker. (Fig. 42.)

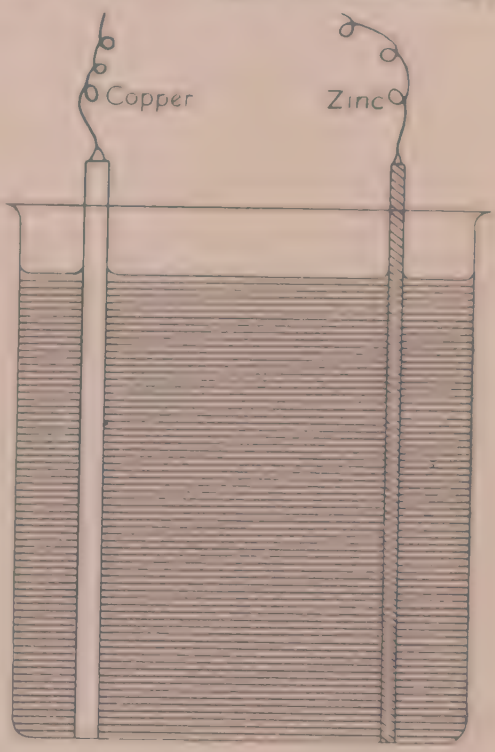

FIG. 42.

Make a summary of the results:

Dilute Sulphuric Acid with—

(*a*) Copper alone → no apparent action.

(*b*) Zinc alone → Hydrogen gas evolved from the surface of the Zinc.

(*c*) Copper and Zinc, separated →

(*d*) Copper and Zinc, in contact →

(*e*) Copper and Zinc, joined by wire →

Experiment 81. Weigh two small pieces of foil—Copper and Zinc—and put them together into a test-tube containing dilute Sulphuric Acid. Test the gas evolved. When the chemical action ceases, remove the Copper foil, wash, dry, and weigh it. Has its weight changed? What has become of the Zinc?

Zinc + Hydrogen sulphate → Zinc sulphate + Hydrogen.
Copper + Hydrogen sulphate → no action.
(dilute)

Three points to note in these experiments are:

(1) The Chemical Action really takes place between the acid and the Zinc.

(2) The Hydrogen evolved during the chemical action appears chiefly at the Copper plate, when the two metals are in contact in the acid or joined by a wire outside.

(3) The connecting wire has a direct influence on what is taking place in the beaker.

A jar containing dilute Sulphuric Acid, in which are immersed a Copper plate and a Zinc plate, joined by a wire, is called a **Simple Cell**.

In diagrams a simple cell is usually represented by two lines—a long line for the Copper and a short line for the Zinc.

Galvani, 1737–98, a doctor of Bologna, had observed in 1786 that the leg of a freshly killed frog, which was hanging by a copper hook from an iron railing, twitched every time the wind carried it against the railing. This doctor's name has passed into our language, and we speak of the apparatus used at bazaars and fairs 'to give electric shocks', or by medical men to administer electricity, as a galvanic battery. Look up 'galvanize' in your dictionary.

Alessandro Volta, 1745–1827, professor at Como and Pavia, showed that the explanation was not to be found in the frog, as Dr. Galvani thought, but in the two *dissimilar* metals, Iron and Copper.

Volta in the course of his experiments recorded the effect on the tongue when it is held between two pieces of dissimilar metals, e.g. Copper and Zinc, and the ends of the metal outside the mouth are brought together. Try it with a florin and a penny.

He constructed what was known as **Volta's Pile.** This was a pile of metal disks, Copper and Zinc, arranged alternately, and separated by pieces of flannel moistened with salt water or dilute sulphuric acid. In 1800 he invented the simple cell.

One of the very important units used in electrical measurement is named the **Volt**, in his honour.

Experiment 82. Fit up a simple cell, and place the free ends of the two wires on your tongue, close together, but not touching.

Remove the plates from the cell, rinse with water to remove acid, and again bring the ends of the wires to your tongue.

Record the results.

An **Electric Current** is said to travel along the wire joining the plates in a simple cell, and in the experiment the tongue is acting as a detector.

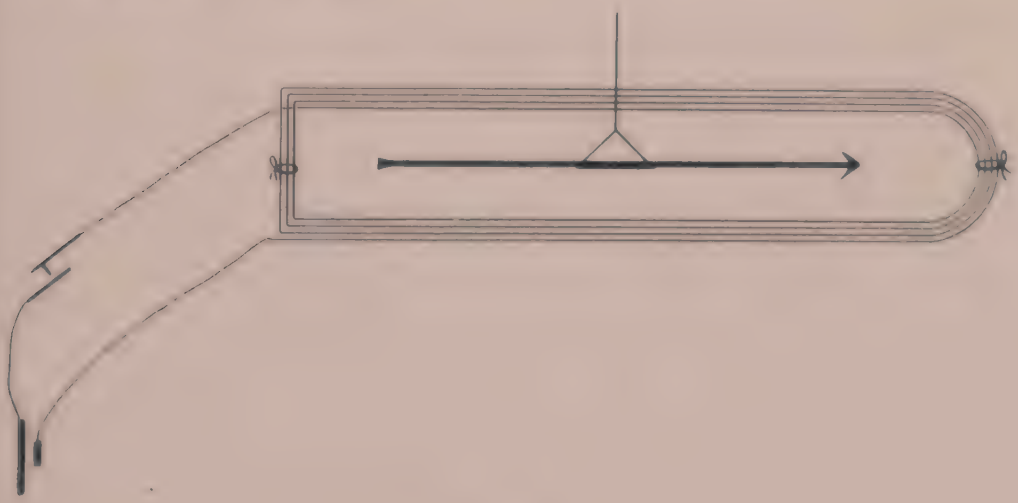

Fig. 43.

Experiment 83. Wind a length of No. 24 d.c.c. copper wire (d.c.c. means double cotton covered) about a dozen times round a book. Tie the coils together at each end. Fit up a magnetic needle (see Figs. 36 *a*, *b*, *c*) and arrange the coil round it as shown in Fig. 43. Connect the wires to a simple cell, and watch the effect on the needle. Record.

The Danish scientist, Oersted, 1777–1851, discovered the

magnetic effect of an electric current, in 1820, by holding a wire conveying the current above a magnetic needle.

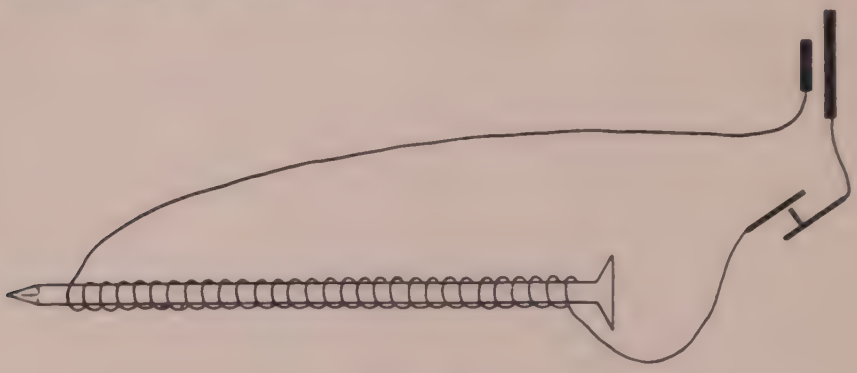

Fig. 44.

Experiment 84. Wind about thirty turns of No. 24 d.c.c. copper wire closely round an iron nail or strip of soft iron without overlapping. (Fig. 44.) Connect the coil to a simple cell, and bring the ends of the nail near the poles of a magnetic needle. Test the nail for magnetism in any other way you know. Record the results.

Make and break the contact, and note the effect on the needle. (Reserve the soft iron and coil for a later experiment.)

Ampère, 1775–1836, a professor in Paris, worked on the discovery made by Oersted, and the results of his *electromagnetic* experiments were published in 1823. He was the first to show that a soft iron core within a coil of wire, along which an electric current was passing, acquired the properties of a magnet.

A piece of soft iron round which is wound a coil of wire conveying an electric current becomes a magnet, and ceases to be a magnet when the current ceases. Such a magnet is termed an **Electromagnet**.

If steel is used instead of *soft iron*, a permanent magnet results.

The simple cell has one serious defect. The Hydrogen that collects on the Copper plate soon causes the strength of the cell to fall off, and many experimenters set themselves to find a remedy.

A very tedious method is to brush the bubbles of Hydrogen away as fast as they appear. Chemical methods, however, are much more satisfactory and, as a rule, depend upon the fact indicated by

Hydrogen + Oxygen → Water.

Potassium bichromate, an orange-yellow crystalline substance, is a powerful oxidizing agent (p. 20).

On adding a solution of potassium bichromate to the Sulphuric Acid of the simple cell, the Hydrogen is oxidized to water as fast as it is formed, and so does not coat the Copper plate.

Experiment 85. Try the effect of adding potassium bichromate solution to the simple cell. Watch the Copper plate. Record.

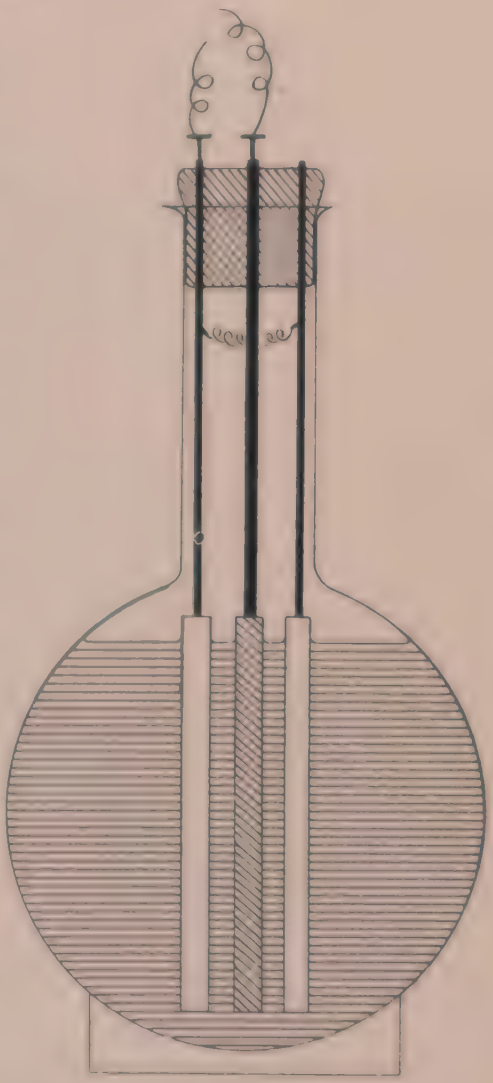

Fig. 45.

In Bichromate cells two Carbon plates—one on each side of the Zinc plate—are used instead of Copper for the sake of economy. (Fig. 45.) The Carbon plates are connected by a wire. The Zinc plate is attached to a rod by which it can be removed from the acid when the cell is not in use.

The most common type of cell is the Leclanché Cell. In it the Hydrogen is oxidized to water by manganese dioxide. The oxidation takes place slowly, so that Leclanché cells are only suitable for intermittent use, e. g. electric bells and flash-

lamps. The Leclanché cell in its *wet form* consists of a Carbon plate and a Zinc rod immersed in a strong solution of sal-ammoniac. The Carbon plate usually stands in a porous pot, i. e. an unglazed pot, tightly packed with manganese dioxide

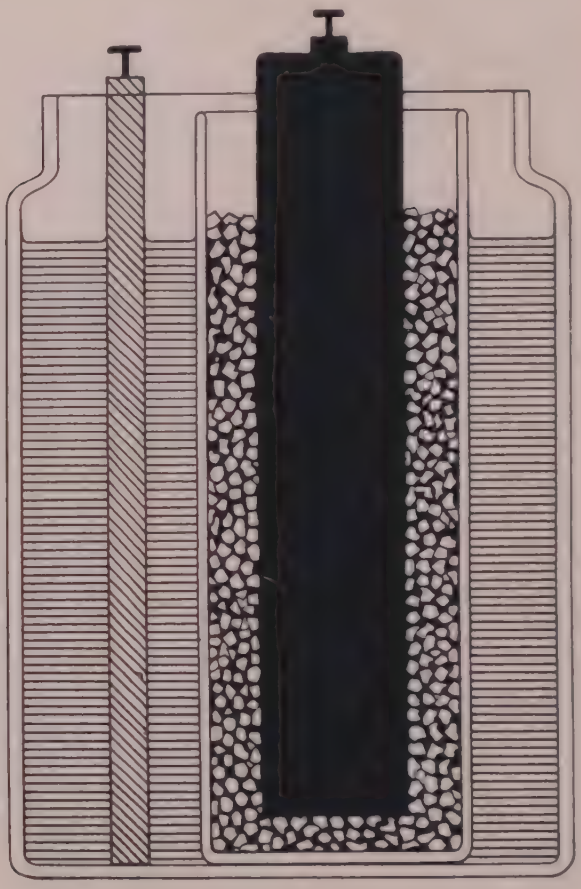

Fig. 46.

and coke. (Fig. 46.) The Zinc rod should be taken out of the solution when the cell is not in use.

The dry form of the Leclanché is in very general use. It is found in almost all pocket flash-lamps, in the High Tension batteries of wireless sets, and, as the source of electricity, for electric bells.

Experiment 86. Obtain an old flash-lamp battery which, as a rule, consists of three dry cells, and pull it to pieces. It is only 'dry' in the sense that it contains nothing spillable. The

solution of sal-ammoniac is frequently made into a paste with plaster of Paris.

Compare the cell with a Leclanché Cell.

When working with cells it is advisable to fit a switch in the circuit so that the current can be cut off when it is no longer required.

In Fig. 47 two forms of switch are shown.

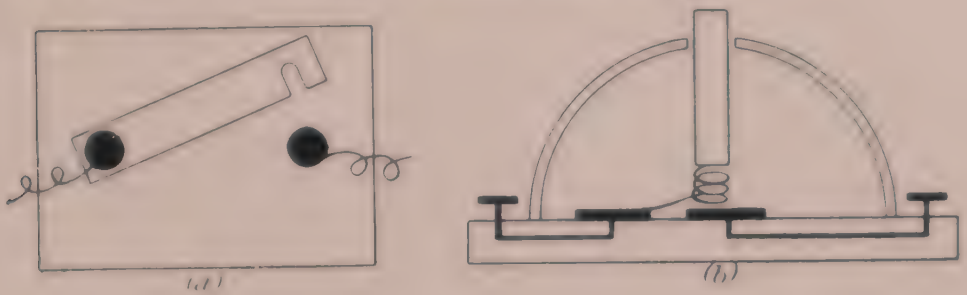

Fig. 47.

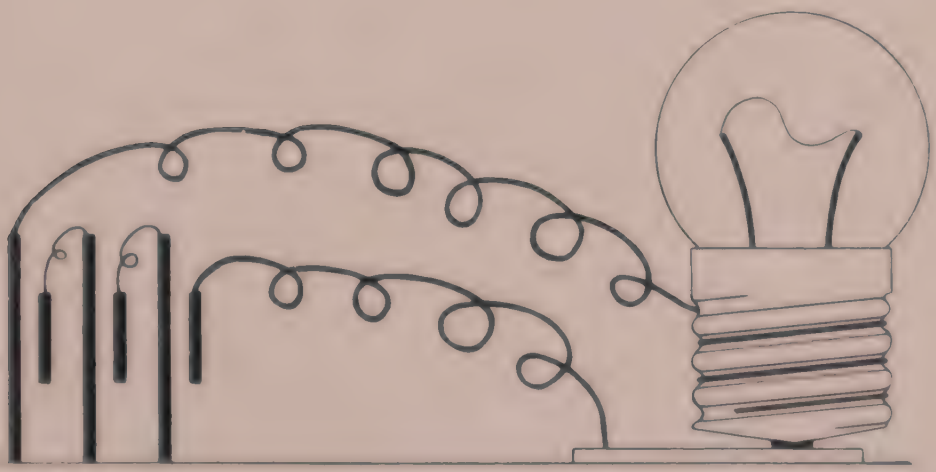

Fig. 48.

(*a*) Two binding screws are fitted to a piece of wood (2 in. square). Contact is made by means of a brass strip.

(*b*) An ordinary bell-push. In this form the current flows only when the button is depressed.

Experiment 87. Take a small electric light bulb such as is used for flash-lamps. Wind the bared end of d.c.c. copper wire tightly round the screw and bind it firmly with a piece of tape.

Attach the other end of the wire to a battery of three cells joined *in series*, i.e. the Copper or Carbon of one cell is connected to the Zinc of its neighbour. Fit up the apparatus as in Fig. 48.

Gather substances commonly employed where electric bells, wireless, &c., are in use, such as copper, wood, brass, rubber, vulcanite, glass, paper, solder, silk, sealing-wax, cotton, and india-rubber.

Place the end of the electric light bulb on a coin, and touch the coin with the end of the other wire. If the apparatus is correctly set up, the lamp will light.

Repeat with the materials gathered.

What does it tell you about these materials (*a*) when the lamp lights, (*b*) when it does not?

Draw up two columns in your note-book, one for conductors and one for insulators. (Lat. *insula*, an island.)

Conductors.	*Insulators.*
Copper.	Glass.

In the light of what you have learned, examine any electrical apparatus you meet, and look out for conductors and insulators.

Experiment 88. Connect two or three cells in series.

Take a short length of very thin bare copper wire and make it part of the circuit by winding one end of it round the bare end of the wire from one terminal and attaching the other end to the other terminal. Record.

Remove the thin copper wire and insert in its place a piece of thin fuse wire. Record.

The Chemical Action taking place in a cell provides Electricity. The question 'What is Electricity?' cannot yet be answered; but, although little is known about Electricity, a great deal is known about its effects.

When the circuit is closed, the wire conveying the electric current exhibits certain properties, e. g.

(1) It may show a physiological effect. (Exp. 82.)

Doctors make use of this in the treatment of nervous diseases.

(2) It may show a magnetic effect. (Exp. 83.)

The magnetic effect is used in many ways. Huge pieces of machinery are lifted by electric cranes. A soft iron core is wound with many turns of wire. This is placed in position, the circuit is completed by means of a switch, and the core becomes a powerful magnet. When the machinery has been removed to the place it is to occupy, the circuit is broken and the core, no longer a magnet, ceases to attract its load.

(3) It may become hot. (Exp. 88.)

Electric fires, ovens, smoothing irons, water heaters are much used.

A very thin wire of metal with a high melting-point, e.g. Tantalum, M.-P. 2850° C., Tungsten, M.-P. 3270° C., raised to a white heat, provides electric light.

Experiment 89. Fit up a magnetic needle and hold, immediately above it, a wire forming part of an electric circuit. Note how the N-pole of the needle is deflected.

Reverse the direction of the electric current, i.e. change the wires at the cell so that the wire that was attached to the Zinc plate is now connected to the Carbon (or Copper) plate and *vice versa*.

Note the deflection of the N-pole of the needle again.

In speaking of an electric current obtained from a cell or cells, the terminal on the Carbon or Copper plate is named the *positive* terminal (+) and the terminal on the Zinc plate is termed the *negative* terminal (−). The current is said to flow from the positive to the negative terminal (+ to −). (Fig. 49.)

In Experiment 89 the deflection of the N-pole of the needle was distinctly shown to depend upon the direction of the current, so that, if the direction of the current is known, the behaviour of the needle can be foretold ; also, if the behaviour

of the needle is known, the direction of the current can be determined.

Ampere's Rule is a very convenient way of remembering the effect of an electric current on a magnetic needle.

'Imagine yourself swimming in the wire with the current and with your face directed towards the magnetic needle; the N-pole is deflected towards the left hand. Notice that when the wire passes below the needle you must imagine yourself

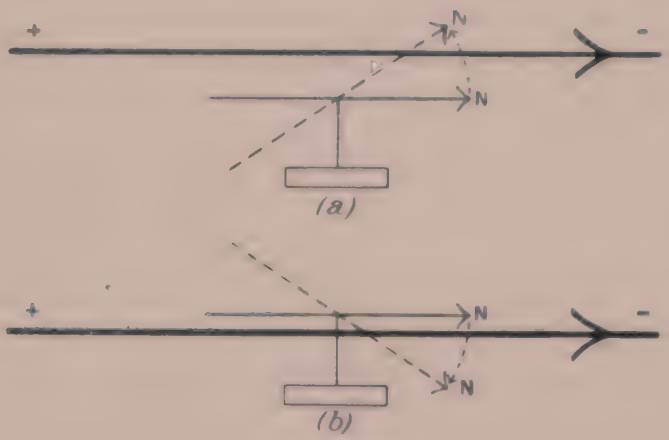

FIG. 49.

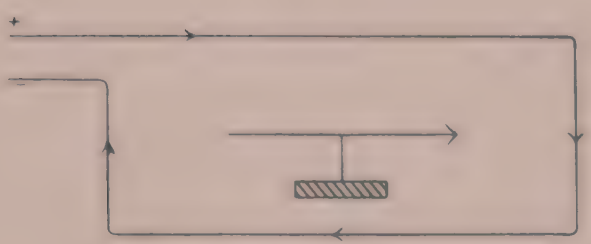

FIG. 50.

swimming on your back, or your face would not be directed *towards* the needle.' Fig. 49 *b*.

Experiment 90. Set up a magnetic needle and hold a wire conveying an electric current

(*a*) above the needle, with current flowing in S → N direction;
(*b*) above the needle, with current flowing in N → S direction;
(*c*) below the needle, with current flowing in S → N direction;
(*d*) below the needle, with current flowing in N → S direction.

Show by diagrams similar to Fig. 49 (*a*) (*b*) the deflection of the needle in each case.

Test the Swimming Rule.

Look at Figs. 50, 51 and consider the effect when a wire is taken completely round the needle. Will the deflection be greater or less than that produced when the current passes the

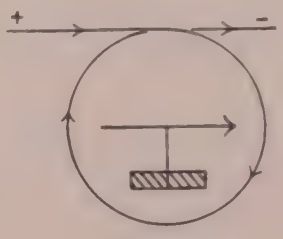

FIG. 51.

needle once? Use the Swimming Rule to answer the question and when you have made up your mind, test your answer by experiment.

Scientific Instruments and Machines, in which electro-magnets are used, employ coils of wire to increase the magnetic effect.

An electric current passed through a chemical compound, water, for instance, as in Experiment 15, may break it up into its elements.

Aluminium is obtained from its ore by electrolysis. This process has made it possible to produce Aluminium so cheaply that many kitchen utensils are made of it.

Bunsen first prepared Aluminium by electrolysis in 1854. In 1855 1 lb. of aluminium was worth £50; to-day that sum would purchase a ton. Copper is obtained in a very pure state by electrolysis.

Electroplating.—Spoons, forks, teapots, cycle hubs, cycle handle-bars, and many other common articles are 'plated'. Often on plated articles used at meal times we see E|P|N|S, which stands for Electroplated on Nickel Silver. Nickel is a white metal resembling silver, but its colour is not so good.

On a clean white tablecloth it appears yellowish. However, as Nickel is much cheaper than Silver, articles are made of it and then coated with real Silver.

Experiment 91. Dissolve crystals of copper sulphate in water slightly acidulated with Sulphuric Acid.

Take two strips of Copper foil (3 in. × ½ in.) and fasten them by drawing-pins to a piece of wood, so that, when the wood is resting across a small beaker, the two strips hang parallel to one another, about an inch apart. Weigh the strip.

Attach to one strip the wire from the + terminal, and to the other the wire from the − terminal, of two or three cells joined *in series*. Mark the wood near the strips + and − respectively. Pour the solution of copper sulphate into the beaker, and watch the strips carefully. In a few minutes weigh the strips.

In this experiment one strip has become thicker at the expense of the other.

The current enters the *electrolytic cell* by the + strip and leaves by the − strip. The + strip loses weight at exactly the same rate as the − strip gains weight.

Michael Faraday (1791–1867), the son of a blacksmith, became one of the most famous experimenters. In 1833 he devoted his attention to electrolysis. To the strip at which the current enters the cell he gave the name **Anode** (Greek *ana*, up, *odos*, the way), and to the strip at which the current leaves the cell he gave the name **Kathode** (Greek *kata*, down).

Using these words, write out a statement describing what happens when an electric current passes from one strip of Copper to another through a solution of copper sulphate.

In electroplating with silver, the anode is a bar of silver, while the kathode consists of the articles to be plated. These are immersed in a bath containing a solution of silver cyanide. The strength of the solution remains unchanged—the metal, in effect, is merely taken from the bar and deposited on the articles.

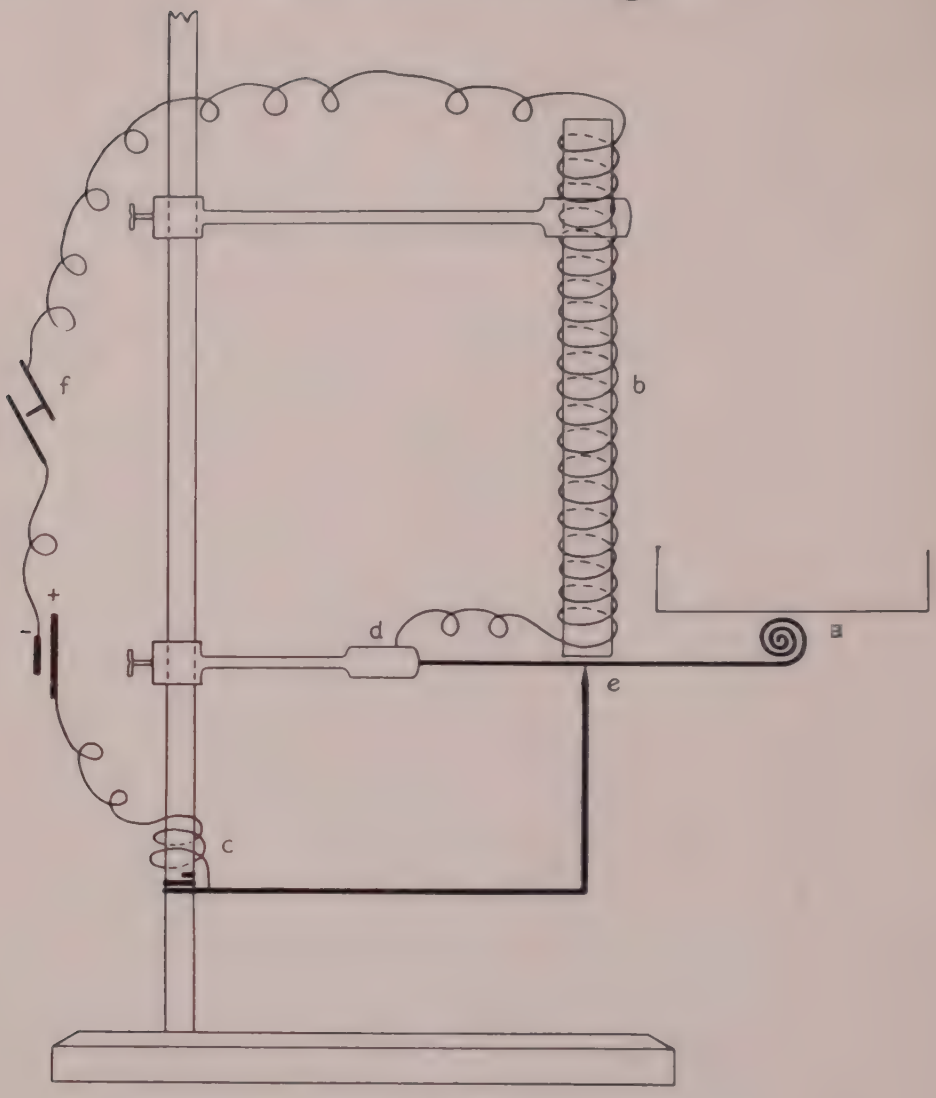

Fig. 52.

The Electric Bell.

Experiment 92 (Fig. 52). Cut a strip of soft iron (8 in. × $\frac{1}{2}$ in.) and roll the end to form a knob (*a*). Wind a close coil of d.c.c. wire round a soft iron rod—a large nail will do (*b*). To the + terminal of a dry cell attach about 9 in. of stout d.c.c. wire (No. 16). Coil it tightly round the rod of the retort stand for support (*c*), and arrange it so that the end points upwards (*e*). Wind one end of the wire from the coil on the soft iron rod round the end of the soft iron strip (*a*), and fix this tightly to a clamp on the retort stand (*d*) so that it just rests on the tip of the wire from the + terminal (*e*). Secure the soft ron rod by another clamp so that its end is just above the strip ($\frac{1}{8}$ in.). Take the other

end of the coil to the −terminal of the cell through a switch (*f*). A cocoa-tin lid above (*a*) will make the 'bell' complete.

Copy the figure, and indicate by arrows the path of the current, when the 'push' is depressed.

You should now be able to explain the working of an electric bell.

It will be interesting at this stage to arrange the discoverers, whose work we have studied, in chronological order. Try to recall, as you read down the list, the discoveries associated with each name.

Note how recent the advance of scientific knowledge is. So long as men tried to find out about Nature and her ways without experimenting little or no progress was made. As soon as they began to experiment progress was rapid; one discovery quickly led to another. When we consider the many modern uses of electricity, it is difficult to realize that the first name associated with the electric current is Galvani (1737–98).

Aristotle 384–322 B. C.
Archimedes 287–212 B. C.
Columbus 1451–1506 A. D.
Guericke 1602–1686.
Torricelli 1608–1647.
Pascal 1623–1662.
Boyle 1627–1691.
Newton 1642–1727.
Stahl 1660–1734.
Fahrenheit 1686–1736.
Celsius 1701–1744.
Franklin 1706–1790.
Black 1728–1799.
Cavendish 1731–1810.
Priestley 1733–1804.
Galvani 1737–1798.
Lavoisier 1743–1794.
Scheele 1742–1786.
Volta 1745–1827.
Proust 1754–1826.
Ampère 1775–1836.
Oersted 1777–1851.
Davy 1778–1829.
Hope 1786–1844.
Faraday 1791–1867.
Liebig 1803–1873.
Bunsen 1811–1899.
Joule 1818–1899.
Kelvin 1824–1907.

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